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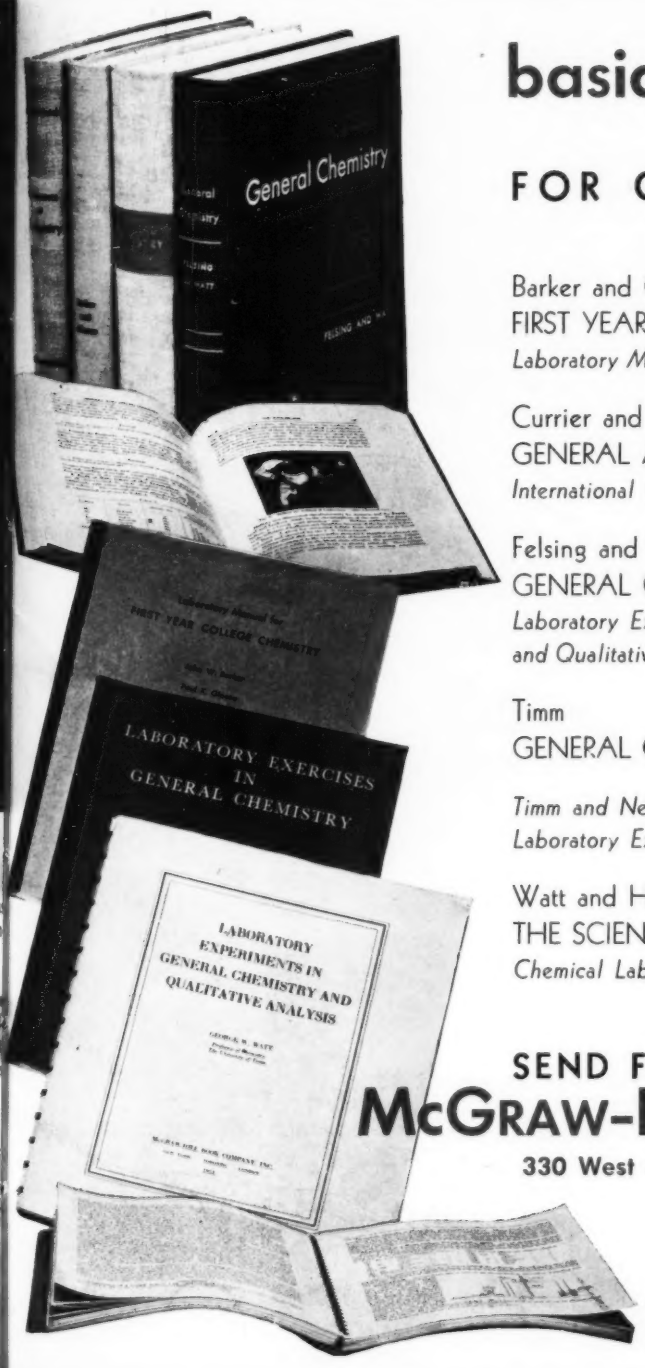
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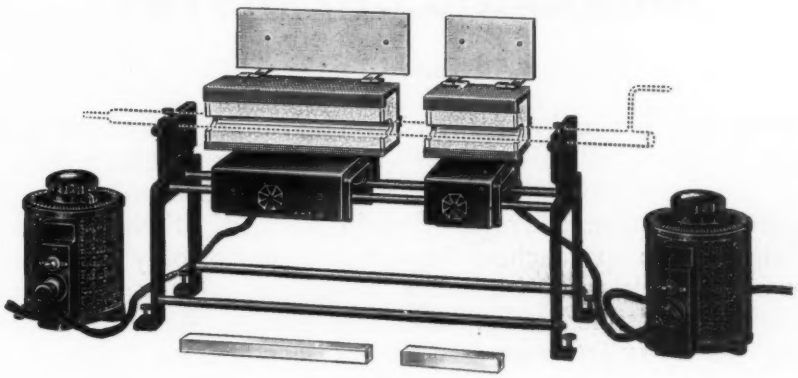

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Unfamiliar Oxidation States: Recent Progress, *Jacob Kleinberg* 324

A Chemistry Identification Question, *Will S. DeLoach* 330

Lemuel F. Smith, 1873–1951, *John R. Sampey* 331

Nuclear Chemistry in the Curricula of American Colleges and Universities, *Russel R. Williams, Jr., and William H. Hamill* 332

Wilhelm Klemm, *Ralph E. Oesper* 336

Coulomb's Law and the Quantitative Interpretation of Reaction Rates, *Edward S. Amis* 337

The Value of the History and Philosophy of Science in the Training of Graduate Students in Chemistry, *Clara de Milt* 340

Language Teaching for Students in Science and Technology, *J. W. Perry* 345

The Boston College Chemical Society, *Albert F. McGuinn, S. J.* 347

Practical Precision for Chemistry Students, *Walter R. Carmody* 349

Jan Ingenhousz, *H. S. van Klooster* 353

The Subject Versus Special Student Interests, *Frederic B. Dutton* 355

The Long Form of the Periodic Table, *George Glockler and A. I. Popov* 358

Proton Waves, *R. S. Bradley* 359

Qualitative Adequacy of Phlogiston, *John Howe Scott* 360

A Simple Procedure for Preparing Pinacol Hydrate, *Joseph E. Weber and Arlo D. Boggs* 363

The Thermomatic Steamer, *Ray Wendland* 364

The 4 in 1 Chem-Mixer, *Ray Wendland* 365

Fragments of Chemistry, *Charles A. Kraus* 366

Proceedings of the Pacific Southwest Association of Chemistry Teachers
Salesmanship in Chemistry Teaching, *J. Rae Schwenck* 368

Report of the New England Association of Chemistry Teachers
Official Business 371

Editor's Outlook 323
Out of the Editor's Basket 25
Letters to the Editor 372
Recent Books 373
Frontispiece—Wilhelm Klemm 21
Advertisers' Index 33

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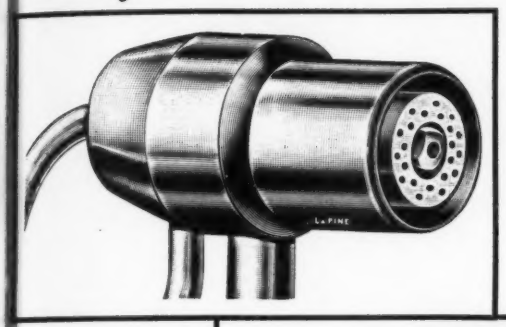

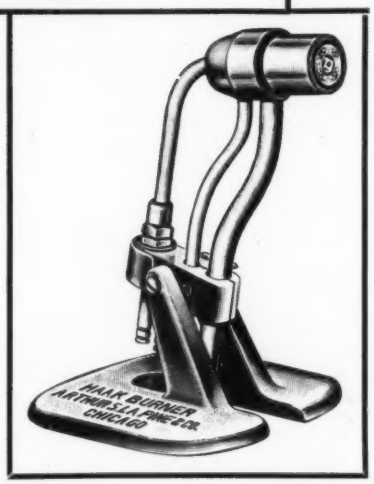

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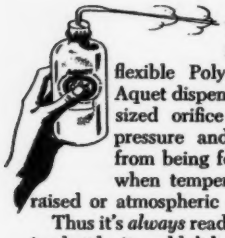

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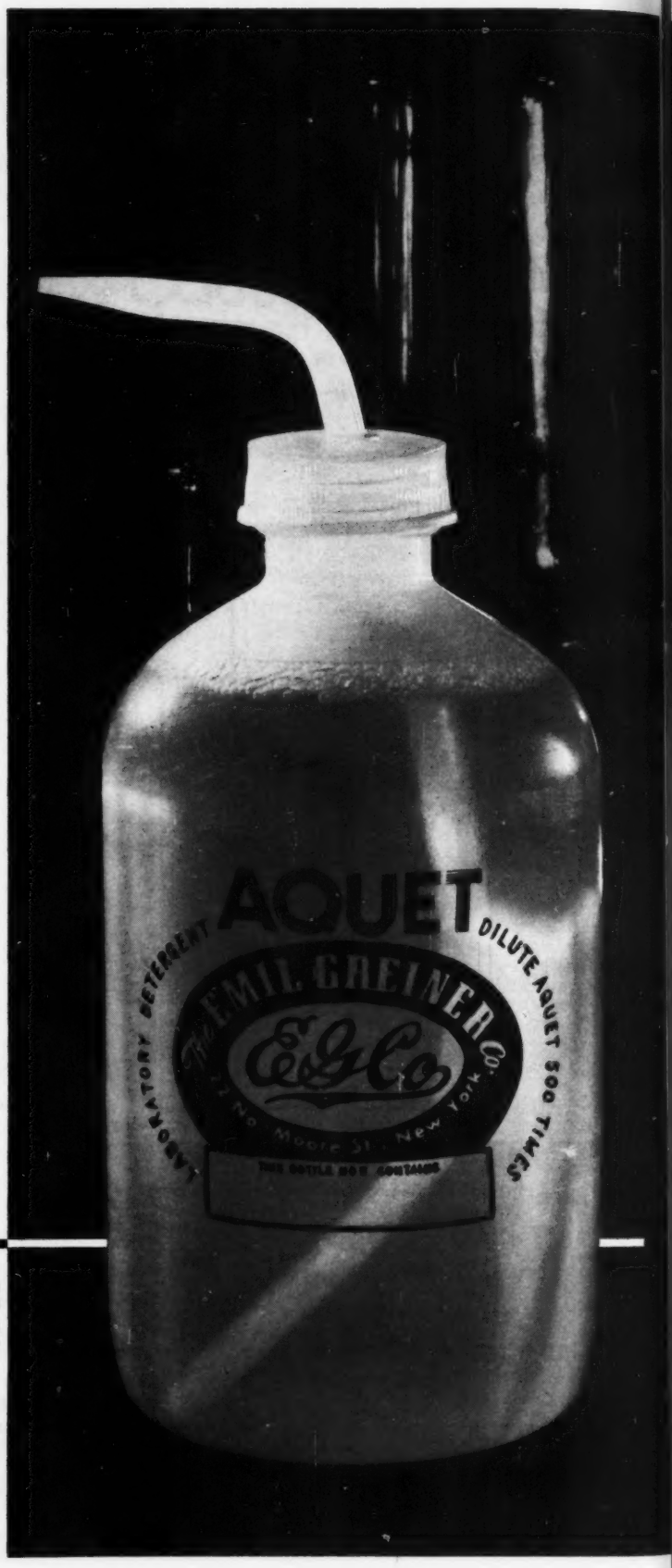

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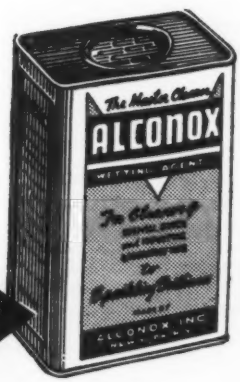

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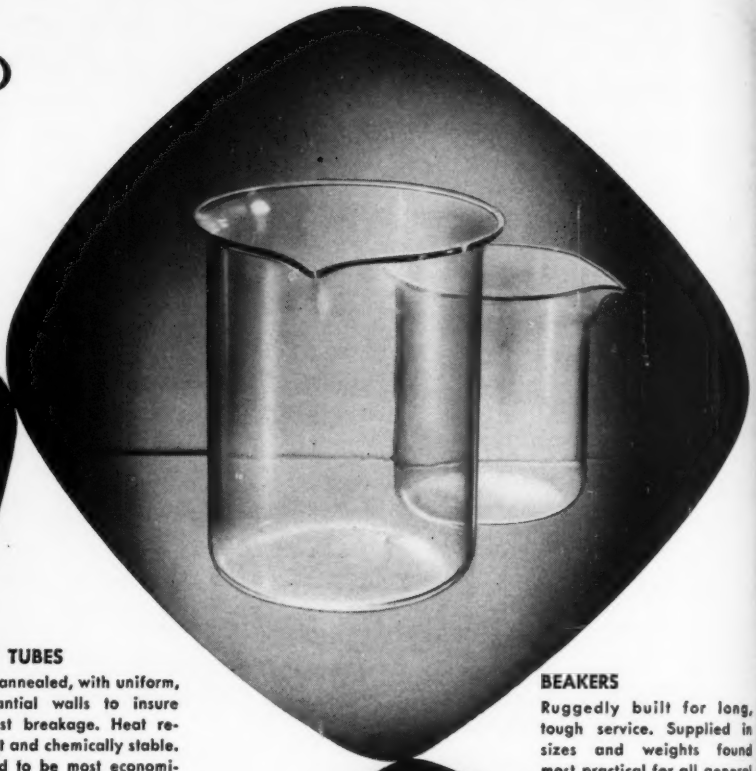

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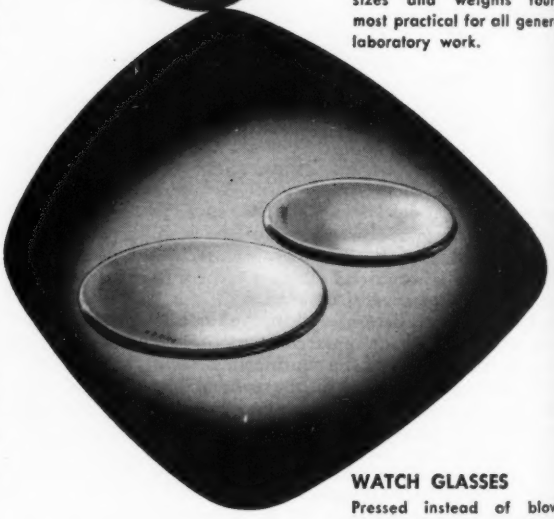

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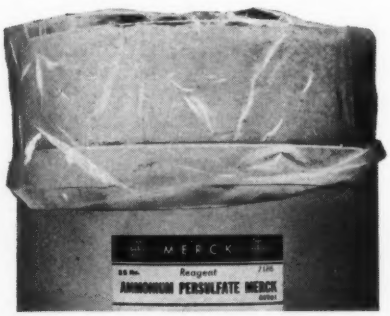

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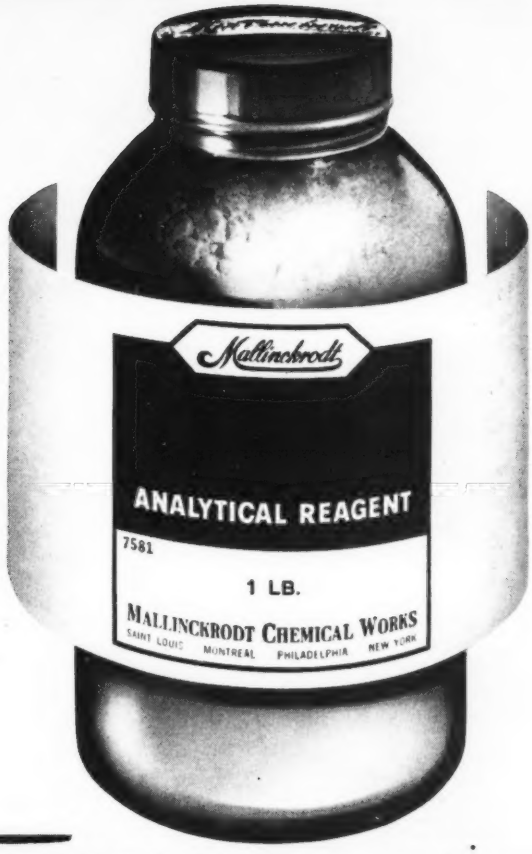

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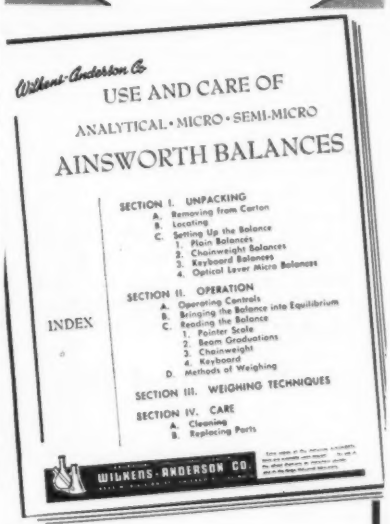

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Type G-1 Mueller Bridge (Excellent Precision) Accessories needed: Same as for Mueller Type G-2	8067	Two ratio arms, 500 Ω each, adjustable to equality. Rheostat range 0 to 81.111 Ω in steps of 0.0001 Ω. Three shunted decades giving steps of 0.0001, 0.001, 0.01 Ω; two decades at end of ratio arms of 0.1 and 1 Ω resistors, and binding posts on end of rheostat for connection of 0, 10, 20, 25, 25.5, 30, 40, 50, 60 or 70-Ω resistors.	±0.02% or a few ten thousandths of an ohm whichever is greater.
Resistance-Thermometer Bridge (Moderate Precision) Accessories needed: Thermohm Resistance Thermometer; Galvanometer	8063	Two ratio arms, 130 Ω nominal. Rheostat range 0 to 200.1 Ω. Two dial decades 9 (1 + 10) Ω, and a 100-Ω resistor removable by short-circuiting link, plus adjustable slidewire of 1.1 Ω.	±0.005 Ω up to 10 Ω ±0.05% above 10 Ω
Portable Temperature Bridge Accessory needed: Thermohm Resistance Thermometer	8062	This instrument is No. 8063 in portable form; all electrical characteristics are the same.	
Portable Temp. Indicator Accessory needed: Thermohm Resistance Thermometer	8015	Range As specified, for type of Thermohm specified	±0.3% of range
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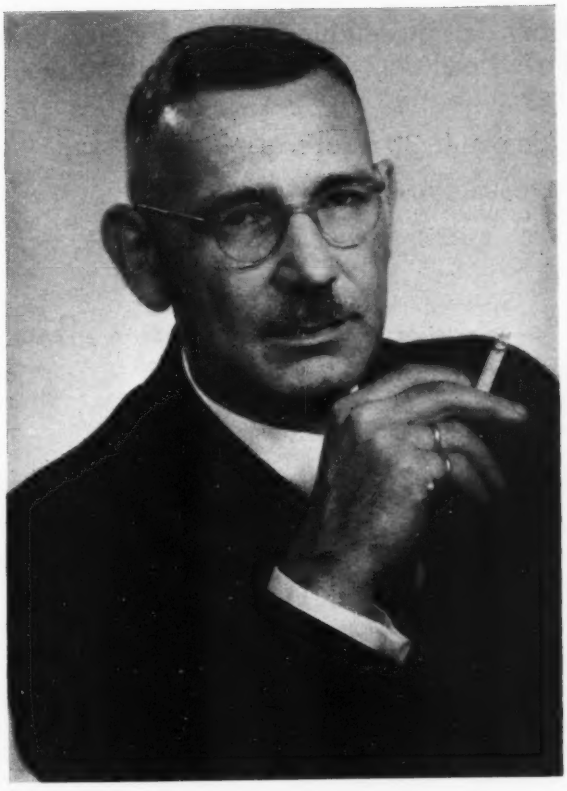

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(See page 336)

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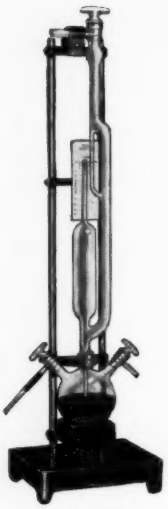

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Editor's Outlook

WHILE running through an accumulation of material, such as piles up on everyone's desk (I hope), I ran across an item, now a year old, which is worth bringing to light. The author, Dr. Webster N. Jones, describes professional training in chemistry at the Carnegie Institute of Technology.[1] I quote from his account:

> The Chemistry Department, working with the other departments in the College of Engineering and Science, has gradually evolved a plan of professional education—one we are proud of because it is working! The aim of professional education at Carnegie is to help the student learn to think; and to equip him to go on learning after graduation and to grow throughout his life in professional and personal stature and in usefulness as a citizen. Carnegie does not seek to *train* students to *be* professional practitioners at graduation, but rather to *educate* them so that they will *become* professional men of full stature. Instruction is therefore planned to help each student acquire the following:
>
> (1) Thorough and integrated understanding of fundamental knowledge and a familiarity with the basic techniques in chemistry and related fields, and ability to use this knowledge and these basic techniques.
>
> (2) Genuine competence in the orderly way of analytical thinking which scientists and other professional men have always used in reaching sound creative conclusions, to the end that after graduation the student can, by such thinking, reach his own decisions in his professional work as a chemist and as a citizen.
>
> (3) Ability to learn for himself with scholarly orderliness, so that after graduation he will be able to grow in wisdom and keep abreast of the changing knowledge and problems of his profession and the society in which he lives.
>
> (4) The philosophical outlook and breadth of knowledge which will enable him to recognize and deal with the human, economic, and social problems of his professional work.

An institution may well be proud of its program if it successfully *educates* its students to *become* professional men. No doubt many institutions aim at this objective, consciously or unconsciously. If in this case the program actually works, as the author indicates, someone deserves commendation.

The methods used in accomplishing this are outlined in the original and cannot be gone into here, but it will be interesting to note the tests which are applied in judging the result. The author says:

> If an affirmative answer can be given to the following questions about a chemist, he is a man of real professional stature.
>
> (1) Does he recognize the real nature of the problem with which he is confronted, and can he state the problem concisely and clearly?
>
> (2) Does he recognize the whole problem including its human, social, and economic elements?
>
> (3) Can he formulate a plan for the solution of the problem and execute the plan?
>
> (4) Can he exercise judgment concerning the general reasonableness or validity of his solution?
>
> (5) Has he the will and the capacity to learn from each experience?
>
> (6) Does he accept his responsibilities as a citizen of his community and the nation?

[1] *The Chemist*, **28**, 248 (June, 1951).

UNFAMILIAR OXIDATION STATES: RECENT PROGRESS

JACOB KLEINBERG
University of Kansas, Lawrence, Kansas

AN IMPORTANT general field of current research in inorganic chemistry is the study of the preparation and characterization of compounds which exhibit elements in their less well-known oxidation states. Recently (*1*, *2*) the author has surveyed in a qualitative manner various aspects of the chemistry of unfamiliar oxidation states of well-known elements,[1] with particular emphasis on the stabilization of such states. The present communication adapts to the pattern of the previous discussions work in this general area which has appeared in approximately the past two years.

ALUMINUM GROUP METALS

Although the existence of aluminum (I) halides in the vapor state has been well-established (*3*, *4*, *5*), only recently has unequivocal evidence been obtained for the formation of lower states of aluminum in liquid ammonia solution and in the solid state. An outstanding achievement along this line has been the preparation of solid aluminum (I) iodide by the use of an electrodeless discharge produced in the vapor of the triiodide at a temperature below 50° (*6*).[2] Such a discharge yields vapors of elementary iodine and a solid material on the walls of the main discharge tube. Exhaustive extraction of the solid with anhydrous benzene, followed by vacuum drying, yields a buff-colored substance, the chemical properties and powder X-ray diffraction pattern of which are consistent with those to be expected of a mixture of metallic aluminum and a monoiodide, $(AlI)_n$. The empirical composition of the best product obtained corresponded to $Al_{1.22}I$.

Evidence for the formation of lower oxidation states of aluminum in liquid ammonia has come from electrochemical studies. Curves obtained from potentiometric titrations of liquid ammonia solutions of aluminum (III) iodide with similar solutions of potassium show *two* distinct end points the first corresponding to the production of Al (II) and the second to Al (I) (*7*). In the anodic oxidation of aluminum in nitrate media in liquid ammonia, the current efficiencies calculated on the assumption of the dissolution of the metal as the tripositive cation are considerably greater than 100 per cent (*8*). In simple nitrate media, apparent initial valence numbers in the neighborhood of 2.7 are found, with little dependence of this value upon the nature of the cation. In nitrate-chloride or nitrate-bromide electrolytes, however, considerably lower valence numbers are obtained; the value varies in a regular manner with the ratio of concentrations of nitrate and halide. In mixtures containing about two mols of nitrate to one of bromide, a minimum valence number of 1.5 is attained: this figure corresponds, at least formally, to a ratio of three unipositive aluminum ions to one tripositive ion. The results show no dependence upon anodic current density, nor upon total concentration of the solutes, so long as the nitrate-halide ratio remains constant. No compound of lower valent aluminum could be isolated from the product of electrolysis. However, nitrate is reduced to elementary nitrogen in quantity equivalent to the amount of Al (I) oxidized to Al (III).

The anodic oxidation of gallium, indium, and thallium in a number of electrolytes in liquid ammonia gives results which vary widely from element to element, as well as from those described above for aluminum (*9*). Gallium is usually oxidized to an initial mean valence number between 2 and 3. With ammonium nitrate-ammonium bromide mixtures as electrolyte, the valence number is a continuous function of the fraction of nitrate in the mixed solute, with a minimum value of 2.0 at about 17 mol per cent nitrate. It is believed that this value may be more plausibly interpreted in terms of the hypothesis of an equimolar mixture of unipositive and tripositive gallium ions, than as indicating the existence of Ga (II). The product of electrolysis in the case of gallium is fairly stable, approximately 90 per cent of the expected reducing power remaining in the anolyte.

With indium as anode, initial valence numbers between 2 and 3 are obtained. With nitrate-bromide mixtures as electrolyte, the valence number does not appear to be a simple function of composition. The lower valent indium formed by electrolysis in ammonium bromide solution is subject not only to oxidation by the electrolyte, but also to rapid disproportionation to the free metal and the tripositive ion.

Anodic oxidation of thallium in ammonia solutions of nitrates yields thallium (I) ion exclusively. In ammonium bromide or potassium amide solutions the anode product contains thallium (III) as well as the uni-

[1] Among the topics treated were the low oxidation states of the metals of the aluminum group; the superoxides of the alkali metals; certain aspects of the chemistry of the positive halogens; little-known compounds of copper, silver, chromium, manganese, iron, cobalt, and nickel; and the "anomalous" oxidation states of the "rare-earths."

[2] In this connection it should be pointed out that calculations by F. Irmann (*Helv. Chim. Acta*, **33**, 1449 (1950)) indicate that solid aluminum(I) halides are only slightly less stable than the common trihalides and should be capable of independent existence.

positive ion. Both ions are stable throughout electrolysis.

SUPEROXIDES AND OZONIDES

Oxygen exhibits relatively unfamiliar oxidation states in superoxides and ozonides, containing O_2^- and O_3^- ions, respectively. Recent studies have contributed interesting information on such compounds.

Superoxides. Superoxides of all the common alkali metals, except lithium, have been prepared (*2*), and evidence is now available that the latter forms an unstable compound of this type (*10*). The rapid oxidation of lithium in liquid ammonia at −78° yields a bright lemon-yellow solution. Measurements of the absorption spectra of oxidized solutions of lithium, sodium, and potassium show that all are similar and exhibit absorption maxima, reasonably close to the same wave length, 380 mμ. Since sodium and potassium are known to form superoxides upon rapid oxidation in liquid ammonia, the similarity of the absorption spectra is strongly indicative that lithium forms a superoxide which is stable in liquid ammonia at −78°. When the solution is permitted to warm up toward −33°, the yellow color disappears and cannot be restored by cooling.

The crystal structure of sodium superoxide, as determined from powder diffraction patterns, consists of a face-centered cubic lattice with $a = 5.490 \pm 0.005$ Å. (*11*). The O_2^- ions, in contrast to the situation with other superoxides, show disorderly orientation. The bond distance in the O_2^- ion is 1.33 ± 0.06 Å. The magnetic susceptibility of sodium superoxide has been measured at a number of temperatures (*12*). A distinct maximum in susceptibility occurs at −80°, the Bohr magneton number decreasing from 2 at this temperature to 0.9 at −196°. It is proposed that this variation is due to superoxide anions which are almost in contact with each other.

The heat of formation of potassium superoxide ($K(s) + O_2(g) \rightarrow KO_2(s)$) has been calculated from the experimental heat of solution of the compound in dilute sulfuric acid solution between 21° and 23° (*13*). The value obtained, $\Delta H = -67.9 \pm 0.1$ kg.-cal./mol, is in good agreement with that previously reported (*14*).

A method for the determination of alkali metal superoxide in the presence of peroxide has been described (*15*). The success of the method depends upon the fact that treatment of a mixture of such substances with an acetic acid-diethyl phthalate solution quantitatively converts superoxide to peroxide and oxygen:

$$2KO_2 + 2HC_2H_3O_2 \rightarrow 2KC_2H_3O_2 + H_2O_2 + O_2$$

little if any decomposition of peroxide occurs.

Attempts to prepare alkaline earth metal superoxides by treatment of nitrates of these metals with alkali metal superoxides in liquid ammonia, although unsuccessful, have led to the discovery of an unusual reaction (*16*). Barium nitrate and potassium superoxide in formula weight ratios up to 1:4 react to yield a tan material which upon analysis is shown to contain potassium, barium, superoxide, and peroxide ions in ratios corresponding to the composition ($2K^+$, Ba^{++}, $2O_2^-$, O_2^{--}). The stoichiometry of the reaction may be represented by the equation

$$Ba(NO_3)_2 + 4KO_2 \rightarrow K_2Ba(O_2)_2O_2 + 2KNO_3 + O_2$$

Of the various combinations tested, only barium nitrate and potassium superoxide give a mixed superoxide-peroxide. Barium nitrate and sodium superoxide in formula weight ratio of 1:2 react to give chiefly barium peroxide. Similarly, from combinations of strontium nitrate and potassium or sodium superoxide, strontium peroxide is obtained. It is noteworthy that in the one case of mixed oxide formation the cations have practically identical ionic radii: K^+, 1.33 Å.; Ba^{++}, 1.35 Å. The mixed superoxide-peroxide is analogous to the so-called "sesquioxides" Rb_4O_6 and Cs_4O_6, which have been shown to contain both superoxide and peroxide ions (*17*).

Ozonides. Although the nature of the products formed by reaction between ozonized oxygen and alkali metal hydroxides has been for many years the subject of speculation, it is only recently that significant progress has been made on this problem. Ozonization of potassium hydroxide at −10° to −15° and extraction of the product with liquid ammonia at −60° to −70° gives, after evaporation of the solvent, a red-brown material which consists mainly of potassium ozonide, KO_3 (*18*). Magnetic measurements show this compound to have a moment corresponding to one unpaired electron, and preliminary X-ray diffraction studies demonstrate that the O_3^- ion is linear, there being a close similarity between the lattices of KO_3 and the corresponding azide.

Studies on the room temperature ozonization of sodium, potassium, and cesium hydroxides have confirmed the existence of ozonides and have also yielded additional information concerning the nature of the ozonization reaction (*19*). Ozone-treated sodium hydroxide decomposes in water with evolution of oxygen, but does not liberate iodine from acidified potassium iodide solution. No colored material is extractable by means of liquid ammonia from the deeply colored ozonated product. Ozonated potassium and cesium hydroxides also evolve oyxgen upon decomposition in water, but these liberate iodine from acidified iodide solution. Analyses of the gross products of ozonization of sodium and cesium hydroxides give data which are most reasonably explained by assigning the formulas NaO_3 and CsO_3 to the colored, paramagnetic materials. Highly colored products are extractable with liquid ammonia from ozone-treated potassium and cesium hydroxides; evaporation to dryness of the resulting solutions produces red solid substances which decompose violently in water with evolution of oxygen, accompanied by liberation of energy in the form of light. Although analysis of the material extracted from ozonated potassium hydroxide shows that the red solid consists mainly of KO_3, chemical behavior would

seem to indicate that the extract is not a single chemical species. Analysis of the extracted cesium material even more strongly suggests the presence of another substance. The substances extracted from both ozonated potassium and cesium hydroxides undergo thermal decomposition to yield products which are apparently the respective superoxides; the gross material resulting from ozonization of sodium hydroxide undergoes decomposition above 100° to produce oxygen and sodium hydroxide.

A recent investigation of the reaction between ozone and sodium hydroxide at temperatures between −50° and −60° has given results differing from those reported above (*20*). According to the investigators, a dark-red liquid ammonia-extractable material (which is chiefly NaO_3) is formed. This substance decomposes at room temperature to yield oxygen and the superoxide. Low temperature ozonization of rubidium and cesium hydroxides, followed by liquid ammonia extraction, results in the isolation of fair yields of the corresponding ozonides. Attempts to prepare lithium ozonide proved unsuccessful.

The heat of solution of potassium ozonide in dilute sulfuric acid, according to the equation

$$2KO_3 + H_2SO_4(aq.) \rightarrow K_2SO_4(aq.) + H_2O + 5/2\ O_2$$

has been determined calorimetrically, and it has been found that $\Delta H = -32.7$ kg.-cal./mol, and the heat of formation (ΔH) from the elements has been calculated to be -62.1 ± 0.9 kg.-cal./mol. A value of 67 kg.-cal. has been obtained for the electron affinity of ozone (*21*).

COPPER AND SILVER

Tripositive Copper. Stable complexes containing copper(III) bound to the periodate (or tellurate) group may be made by the oxidation in alkaline medium of copper(II) by means of persulfate in the presence of the complexing group (*22*). A periodate complex of the formula $K_7[Cu(IO_6)_2]$ has recently found analytical application as an oxidimetric titrant (*23*, *24*). It has been used for the determination of dextrose in blood and urine, one mol of sugar reacting with eight atoms of tripositive copper in alkaline solution (*23*). An important application is the characterization of proteins; every protein can be assigned a distinctive "percupri" number (*24*). Calcium may be determined by precipitation with an alkaline solution of napthalhydroxamic acid ($C_{10}H_6(CO)_2NOH$), followed by titration of the naphthalhydroxamate with the deep-brown "percupri" complex in hot alkaline medium (*24*). One mol of the calcium compound rapidly converts four atoms of tripositive copper to the dipositive state, and then two more copper atoms are slowly reduced. In the first stage of the reaction, the oxime group is converted to nitrite and naphthalic acid is produced; in the second stage nitrite is oxidized to nitrate. In dilute alkaline medium the "percupri" complex oxidizes arsenite to arsenate, antimonite to antimonate, chromic ion to chromate, cyanide ion to carbonate and nitrate, thiosulfate to sulfate, and a number of sulfides to sulfates (*24*).

At 250° a mixture of potassium and copper (II) chlorides, in the formula weight ratio 3:1, reacts with fluorine to produce a homogeneous blue-green substance of the formula K_3CuF_6 (*25*). This compound is decomposed by water to yield copper (II) fluoride and an unidentified gas.

Di- and Tripositive Silver. Silver (II) oxide, which may be prepared conveniently by the oxidation of silver (I) nitrate with potassium persulfate in alkaline medium (*26*), is beginning to find application as an oxidizing agent in analytical chemistry. It is claimed to be a more efficient oxidizing agent than either persulfate in a solution containing unipositive silver or sodium bismuthate (*27*). As little as 1 p. p. m. of manganese or 1 mg. of cerium (III) ion may be detected by the use of this reagent.

Although dipositive silver possesses one unpaired electron, magnetic measurements on the oxide between −183° and +95° (*28*) show that the oxide is diamagnetic, having a molar susceptibility of -19.1×10^{-6} unit at room temperature. On the other hand, the silver ion in the dark-brown solution resulting from the dissolution of silver (II) oxide in concentrated nitric acid is highly paramagnetic, with an effective Bohr magneton number of 1.76, a value in excellent agreement with that expected for an ion with one unpaired electron (*28*).

Silver (II) oxide is reduced in dilute nitric acid solution with the evolution of oxygen. However, the dipositive silver may be stabilized by the presence of α,α-dipyridyl in the nitric acid solution (*29*). Addition of a large excess of cold, saturated potassium perchlorate solution to such a solution results in the precipitation of red crystals of α,α-dipyridylsilver (II) perchlorate, $[Ag(dipy)_2](ClO_4)_2$. The stabilizing effect on dipositive silver of coordination with dipyridyl is evident from the value of the reduction potential of the complex ion, which has been reported to be 1.453 volts (*30*), whereas the formal reduction potential for the reaction

$$Ag(II) + e^- \rightleftharpoons Ag(I)$$

in 4 *M* perchloric acid at 25° has been found to be 1.9998 volts (*31*).

The only cationic complexes of tripositive silver which have been reported are those containing ethylenedibiguanide, $[C_2H_4(BigH)_2]$, as coordinating agent (*32*). Evidence for the great stability of these complexes in aqueous solutions has been obtained from determinations of the pK values for the following equilibria (*33*):

$$(a)\ Ag[C_2H_4(BigH)_2]^{+3} \rightleftharpoons Ag^{+3} + C_2H_4(BigH)_2$$
$$(b)\ Ag[C_2H_4(BigH)_2]^{+3} + 2H^+ \rightleftharpoons Ag^{+3} + C_2H_4(BigH_2)_2^{+3}$$

The pK values found were: for equilibrium (*a*), 52.3; (*b*) 29.2. As might be expected, the complex ion undergoes extensive hydrolysis in solutions of high pH.

CHROMIUM AND MANGANESE

Dipositive Chromium. Magnetic studies on chromium (II) acetate (and its monohydrate) give evidence for d^2s binding in this compound (*34*). The anhydrous acetate has a susceptibility of 430×10^{-6} c. g. s. unit per mol, whereas the monohydrate has a value of -85×10^{-6} unit. The diamagnetic character of these compounds (the small degree of paramagnetism of the anhydrous salt is believed to be due to the presence of some chromium (III)) demonstrates that the four unpaired *3d* electrons of the normal Cr^{+2} ion are paired and that the three orbitals thus made available are involved in the formation of covalent bonds. Thus the chromium atom presumably is attached to four acetate groups by d^2s binding.

Dipositive chromium is finding analytical use as a reductant. Dissolved oxygen has been determined successfully by titration with chromium (II) ion in acid medium (*35*). Potentiometric titrations of Sb (V), Cu (II), Sn (IV) (*36*), and Bi (III) (*37*) have been investigated. The first three were determined most effectively in solutions 5 *N* or more concentrated in hydrochloric acid at 85°. Antimony is converted to the tripositive and tin to the dipositive state. In sulfuric acid solution Cu (II) is reduced directly to the elemental state, but in hydrochloric acid medium the titration curve shows breaks corresponding to Cu (II)→Cu (I)→Cu°. The second equivalence point is defined clearly only when the solution is at room temperature or below, and disappears when the titration is carried out at 85°. The tin determination is satisfactory in the presence of an equal or smaller amount of antimony and/or copper, but not with larger amounts. Antimony-copper mixtures with ratios from about 100:1 to 0.1:1 can be analyzed accurately. Small amounts of antimony and/or copper can be determined in the presence of large quantities of tin. Tripositive bismuth is converted to the free element by chromium (II); the titration can be successfully carried out in hydrochloric acid solutions of concentration 0.6 to 6 *N*. Bismuth can be determined in the presence of lead and cadmium by this method.

Tetra- and Pentapositive Chromium. Fluorination of a 3:1 molar mixture of potassium chloride–chromium (III) chloride produces, in addition to a small amount of chromium (V) fluoride which sublimes from the mixture, a flesh-colored solid. This solid, which rapidly turns yellow on standing, has been shown by analysis to be potassium hexafluochromate (IV), K_2CrF_6 (*38*).[3] The X-ray pattern of this substance is similar to that of K_2MnF_6. At elevated temperatures the hexafluochromate disproportionates to give chromium (V) fluoride and an unidentified chromium-(III) compound.

Reaction at 130–150° between bromine trifluoride and an equimolar mixture of potassium chloride–chromium (VI) oxide, or between the trifluoride and potassium dichromate, results in the formation of a compound of empirical formula $KCrOF_4$ (*39*). A similar reaction with silver dichromate yields $AgCrOF_4$. The light color of these tetrafluooxychromates appears to preclude the possibility of the presence of chromium in two different oxidation states, and is indicative of the presence of chromium (V). In water, the fluooxychromates undergo immediate hydrolysis with the formation of tri- and hexapositive chromium.

Pentapositive Manganese. An interesting illustration of the phenomenon of *valence inductivity* (*40*) is found in the formation of manganese (V) from lower oxidation states of this element or from manganese (VII) (*41*). The dehydration at 50–150° of an intimate mixture of trisodium phosphate (1 mol) and manganese-(II) acetate tetrahydrate (or potassium permanganate) (0.01 mol), followed by heating at 750–800° in a closed container, results in the formation of the blue trisodium manganate (V), Na_3MnO_4. Replacement of the phosphate in the original mixture by sulfates yields no manganese in an oxidation state greater than four. Manganese (V) is more stable in the absence of free alkali, both at room temperature and at 700–800°, than is manganese (VI). Trisodium manganate (V) is readily soluble in 5 per cent aqueous potassium hydroxide, the manganese (V) ultimately disproportionating to yield manganese (VI) and the dioxide.

IRON, COBALT, AND NICKEL

Hexapositive Iron. Alkali metal ferrates (VI), M_2FeO_4, are most conveniently obtained by the chlorine oxidation of suspensions of hydrous iron (III) oxide in concentrated alkaline solution. The products so prepared are contaminated with potassium chloride and unreacted iron oxide. In view of the potential applications of ferrates in oxidimetry, the preparation of analytically pure material has been the subject of investigation (*42*, *43*). A simple method of purification which reportedly yields potassium ferrate of analytical purity is briefly described (*42*). Potassium ferrate (VI) and chloride are extracted from the iron oxide by means of ice water. The addition of concentrated potassium hydroxide solution to the extract results in the reprecipitation of the ferrate and chloride. The latter is then removed by treatment with pure methanol at −10°, and the residual potassium ferrate is finally washed with cooled ether and dried at 80°. Above 80° potassium ferrate (VI) decomposes with the evolution of oxygen. In alkaline medium the ferrate quantitatively converts chromium (III) to chromium (VI) and arsenite to arsenate (*44*). In acid medium one formula weight of the substance liberates four equivalents of iodine from iodide (*42*).

X-ray diagrams show marked similarities between FeO_4^{--}, SO_4^{--}, and CrO_4^{--} ions (*42*). The potassium salts of these ions form mixed crystals. It is interesting that the stability of potassium ferrate (VI) has

[3] The simple fluoride CrF_4 has been prepared by reaction between fluorine and elementary chromium or its halides at 350–500°. (See v. WARTENBERG, H., *Z. anorg. allgem. Chem.*, **247**, 135 (1951).)

been reported to be enhanced in mixed crystals containing potassium sulfate, but lowered in solutions containing the latter substance.

The magnetic susceptibility of potassium ferrate (VI) has been determined by the Gouy method (*45*). The best value obtained for the effective magnetic moment was 3.06 ± 0.02 Bohr magnetons. This is in excellent agreement with the theoretical value (2.83 Bohr magnetons) expected for hexapositive iron with two unpaired *3d* electrons.

Unipositive Cobalt. There is evidence that unipositive cobalt is formed in the polarographic reduction of cobalt (II) chloride hexahydrate in a solvent melt composed of 30 mol per cent lithium nitrate, 17 mol per cent sodium nitrate, and 53 mol per cent potassium nitrate (m. p. 120°) (*46*). A double wave is obtained; this is attributed to oxidation of cobalt to the tripositive state by the nitrate melt, with subsequent stepwise reduction to the unipositive state and the metal. The plateau of the second wave is 1.3 times as high as the first; theoretical value for the stepwise reduction proposed is 1.5.

Tetrapositive Cobalt. A complex fluoride of the formula K_3CoF_7 has been prepared by reaction at 325–425° between fluorine and a mixture of potassium and cobalt (II) chlorides (formula weight ratio 3:1) (*25*). The compound is relatively stable toward water even at elevated temperatures. X-ray measurements which show that the complex fluoride possesses a cubic structure of the K_3ZrF_7 type confirm the presence of tetrapositive cobalt.

Nickel (0). Replacement of carbon monoxide in nickel tetracarbonyl by other neutral molecules has led to the synthesis of some new compounds containing nickel (0). Reaction between phenyl isocyanide and nickel tetracarbonyl in anhydrous ether gives in several minutes canary-yellow prisms of tetrakisphenylisocyanide nickel, $Ni(C_6H_5NC)_4$ (*47*). The nickel (0) compound is a stable substance (decomposition point 105°), soluble in ether and benzene and insoluble in water. With nickel tetracarbonyl and methyl isocyanide only three carbon monoxide units are replaced and $Ni[CO(CH_3NC)_3]$ is formed.

Complexes containing nickel (0) bound to phosphorus(III) chloride, bromide, or fluoride have also been prepared (*48*, *49*). Tetrakistrichlorophosphine nickel, $Ni(PCl_3)_4$ is formed by treatment of nickel tetracarbonyl with an excess of phosphorus (III) chloride, followed by removal of carbon monoxide by heating (*48*). The pure compound is a pale-yellow solid, which is stable up to 120°; above this temperature blackening and loss of phosphorus (III) chloride occur. It is soluble, although more or less unstable, in organic solvents. It is rapidly attacked by hot acid solutions and slowly by concentrated sodium hydroxide solution. The latter gives a black precipitate and a solution of sodium phosphite. Molecular weight determinations in benzene give a value corresponding to a monomeric state of $Ni(PCl_3)_4$. Magnetic susceptibility determinations at 25° give an apparent molar susceptibility of $+33 \times 10^{-6}$ c. g. s. unit for the nickel in the compound, a value which indicates the absence of unpaired electrons.

Tetrakistribromo- and trifluorophosphine nickel have been prepared in good yield, the former most conveniently by treatment of the trichloro compound with phosphorus (III) bromide at 70°, and the latter by reaction of the trichloro or the tribromo compound with phosphorus (III) fluoride at 100° and 50–100 atmospheres pressure (*49*). $Ni(PF_3)_4$ is the most stable of the halophosphine nickel derivatives. It is very soluble in benzene and toluene, and can be volatilized in steam with little hydrolysis. It boils at 70.7 ± 0.3° at 760 mm. and has a density of 1.80 at 25°.

The interaction of antimony (III) chloride with nickel tetracarbonyl and iron pentacarbonyl yields mixed carbonyltrichlorostibine products; $Ni[(CO)_3SbCl_3]$ from reaction in cyclohexane medium and $Fe[(CO)_3(SbCl_3)_2]$ from benzene, respectively (*50*). Both the nickel and the iron complexes are insoluble in such organic solvents as benzene, pentane, cyclohexane, and carbon tetrachloride. Carbon disulfide reacts to liberate carbon monoxide and dark colored residues. Reaction with alcohols or acetone results in decomposition of the complex, carbon monoxide being evolved and nickel (II) or iron (II) being formed. Water and dilute acids and alkalies also cause rapid decomposition.

Tripositive Nickel. Nickel (II) salts form a series of compounds with the chelating agent *o*-phenylenebidimethylarsine, $o\text{-}C_6H_4(As(CH_3)_2)_2$, of the general formula $[Ni(C_6H_4(As(CH_3)_2)_2]X_2$. These complexes are diamagnetic, and therefore planar with dsp^2 binding. The chloride when refluxed with hydrochloric acid in air is oxidized to the octahedral greenish-yellow $[Ni(C_6H_4(As(CH_3)_2)_2)_2Cl_2]Cl$ (*51*). Proof of the formula shown above arises from conductivity studies, titration of ionic chlorine with silver ion, and magnetic measurements, the latter showing the presence of one unpaired electron in the complex. The exact nature of the binding in the complex is undetermined; it may involve dsp^3 covalent bonds and one ionic bond, or d^2sp^3 bonds and promotion of one electron to the *4d* level of the nickel. The tripositive nickel complex is stable in air, but can be reduced to the complex containing dipositive nickel by such reducing agents as sulfur dioxide or tin (II) chloride. $[Ni(C_6H_4(As(CH_3)_2)_2)_2Cl_2]Cl$ may be converted to compounds of the general formula $Ni(C_6H_4(As(CH_3)_2)_2)_2X_3$ by treatment with the appropriate anion (X = Br^-, NCS^-, I^-). The bromide and thiocyanate complexes are similar to the chloride containing tripositive nickel, whereas the iodide is diamagnetic and thus a nickel (II) compound, possibly containing the I_3^- anion.

Higher Hydroxides of Nickel. Evidence has been obtained for the existence of a number of higher hydroxides of nickel from a calorimetric study of the oxidation of nickel (II) nitrate by means of alkaline potassium persulfate, and from the anodic oxidation of an electrolyte consisting of nickel (II) sulfate and sodium acetate (*52*). Nickel (IV) oxide $NiO_2.xH_2O$, was detected

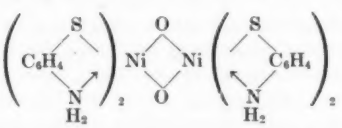
only in the calorimetric experiments and could not be isolated because of its instability. A tripositive nickel hydroxide, designated β-NiOOH, is produced at room temperature by oxidation of dipositive nickel both with persulfate and anodically. A substance of the composition $Ni_3O_2(OH)_4$ is formed both by anodic oxidation at 40–60° and by persulfate oxidation at 100°.

These studies on the higher hydroxides of nickel have led to the conclusion that the primary product formed at the anode during the charging of an Edison cell is $NiO_2.xH_2O$ (*53*). This substance then forms β-NiOOH both by decomposition and by reaction with nickel (II) hydroxide. During the discharge process, decomposition of $NiO_2.xH_2O$ results in a rapid initial fall of potential to that corresponding to the formation of β-NiOOH. On slow discharge a potential drop corresponding to the production of $Ni_3O_2(OH)_4$ is observed.

Tetrapositive Nickel. A black product consisting chiefly of $BaNiO_3$ has been obtained by reaction at 700° between oxygen and an equimolar mixture of barium hydroxide and nickel (II) oxide (*54*). Beginning at 730° (± 10°) at an oxygen pressure of 730 mm. structural changes occur which are associated with the formation of substances with compositions ranging between $Ba_3Ni_2O_8$ and $Ba_2Ni_2O_5$. $BaNiO_3$ is insoluble in alkaline solutions but reacts readily with acids (*e. g.*, hydrochloric acid is oxidized quantitatively to chlorine). It fails to give a peroxide test. The magnetic susceptibility of the compound is very low, the calculated number of unpaired electrons per nickel atom being 0.82 at 300°K.

A red fluocomplex of the composition K_2NiF_6 has been obtained by the fluorination at 275° of a 2:1 molar mixture of potassium and nickel (II) chlorides (*25*). X-ray measurements, which show that the complex exhibits a cubic structure of the K_2PtCl_6 type, substantiate the presence of tetrapositive nickel. The compound differs from the tetrapositive cobalt fluocomplex in that it is decomposed by water with evolution of gas (unidentified) and the formation of a black precipitate (unidentified).

Stable nickel (IV) complexes of organic thio salts have been reported (*55*). The air oxidation of polymeric nickel (II) *o*-aminothiophenolate suspended in strong alkali yields a deep-blue μ-dioxonickel(IV) complex of the constitution



A somewhat similar disulfo compound has been characterized. This substance dissolves in organic solvents to give deep-violet solutions. It is formed, along with nickel tetracarbonyl, when the nickel (II) salt of dithiobenzoic acid is treated with carbon monoxide and SH^- ion.

THE "RARE-EARTHS"

A report (*56*) that praseodymium (V) oxide is formed when the sesquioxide is heated with oxygen under pressure (15 atm.) has not been confirmed. Recent work (*57*) has shown conclusively that praseodymium (IV) oxide is the highest oxide formed when Pr_2O_3-Nd_2O_3 and Pr_2O_3-CeO_2 mixtures are heated at 300° under 50 atm. oxygen pressure. Treatment of the sesquioxides of praseodymium and terbium with atomic oxygen yields oxides containing these elements in the tetrapositive state; no evidence for the formation of any higher oxide than the sesquioxide is obtained with neodymium, samarium, europium, gadolinium, or ytterbium (*58*).

The recent report (*59*) of the existence of praseodymium (IV) and neodymium (IV), presumably as complexes with 8-quinolinol-5-sulfonic acid, has not been substantiated (*60*).

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A CHEMISTRY IDENTIFICATION QUESTION

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GIVEN seven substances studied this semester, identified by Numbers 1, 2, 3, 4, 5, 6, and 7. Nos. 1, 2, 5, and 6 are elements. Nos. 1, 3, and 7 are gases ordinarily. Nos. 2, 5, and 6 are solids ordinarily. Many of them have oxidizing or reducing properties but 1 is a strong oxidant, 5 is a strong reductant, and 4 is a strong oxidant when hot. All seven are produced on an important scale commercially. When 2 is burned in air, 7 is formed. No. 3 dissolves in water to form a basic solution. Nos. 4 and 7 form acidic solutions in water. No. 7 is used in the industrial preparation of 4. All of the gases have odors, and 3 and 7 are colorless. No. 5 reacts vigorously with water, forming a basic solution. No. 6 will burn in air to form a compound that is a good dehydrating agent. The four elements arranged in order of increasing number of valence electrons are: 5, 6, 2, 1. Nos. 5 and 6 must be stored out of contact with air. Nos. 3 and 4 contain one element in common, and 4 and 7 contain two elements in common. Nos. 1 and 7 have industrial use as bleaches. No. 2 has two crystalline forms. Nos. 3 and 4 in water solution are common laboratory reagents. No. 5 has a silvery luster. No. 6 has two allotropic forms, one of which is much more reactive than the other. The reactive form of 6 is quite poisonous. Nos. 3 and 7 have some use in electric refrigerators.

Identify the seven substances by writing their names or formulas in the spaces below.

"Scrambled clues" questions of this sort have been used by the author to add variety to general and organic chemistry tests and examinations, to catch the interest of the better students, and to require more than mere recall of memorized facts and theories. Such questions are not tedious to prepare, can touch on much material, and can be scored rapidly. It is necessary to take care to keep them from being too easy—so that everyone answers them—or so difficult as to be impossible with the knowledge the students are expected to have. The student's difficulties in analyzing the problem can of course be increased by having a larger number of unknowns or by making the clues less obvious. It seems to be a good idea to show relationships between the unknowns as far as possible, but not to require the solution of one to depend completely on another. Time is an important factor since many students are unable to make much progress on this sort of problem if they feel rushed.

One such question was given to a class in general chemistry on the final examination at the end of each

semester, and the table gives the correlation coefficients between the score on this item, in each case, and the total score on the remainder of the examination.

Question	*Correlation coefficient*[a]	*Number of persons*
1	$r_{1,\,t-1} = 0.70$	24
2	$r_{2,\,t-2} = 0.68$	22
1	$r_{1,\,2} = 0.61$	21

[a] $r_{1,\,t-1}$ is the correlation coefficient between item 1 and the total score on the test excluding item 1.

Although these data are based on a rather small number of cases the coefficients—0.70 and 0.68—indicate that this type of question is probably substantially valid. The correlation coefficient between the scores on the two questions is also given for those persons, 21 in number, who took both examinations. This figure—0.61—gives some measure of the reliability of this type of question, there being a considerable time lapse and a difference in content between the two questions.

The answers to the question given at the beginning are: (1) Cl_2, (2) S, (3) NH_3, (4) H_2SO_4, (5) Na, (6) P, and (7) SO_2. The scores actually obtained when this question was given to a class of 22 students ranged from zero (2 persons) to the maximum possible (also 2 persons).

ACKNOWLEDGMENT

The author wishes to thank Dr. Julian C. Stanley, Jr., of the Psychology Department, George Peabody College, for his help in preparing this paper.

LEMUEL F. SMITH, 1873–1951

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THE sudden passing of Professor Lemuel F. Smith on the evening of November 24, 1951, brought to a close the life of one of the inspiring teachers of this generation. During a third of a century as Professor of Chemistry at Kalamazoo College he saw more than 60 of his students complete their studies for the Ph.D. degree in the graduate schools of the nation. Recent research into the origins of American scientists[1] disclosed that Kalamazoo stands third among the institutions of higher learning in the country in the training of scientists.

Dr. Smith continued to teach after he reached retirement age. He spent one year at Amherst, two at Georgetown (Kentucky), three at Furman, one at St. Lawrence, and he was in what he termed his last year at Richmond when his labors were brought quietly to a close as he was listening to his radio just before retiring for the night.

All who knew Dr. Smith enjoyed his refreshing sense of humor. On one occasion the president of the college told him he had just secured a young teacher of religion named Lemuel E. Smith, and asked how he could distinguish them. Dr. Smith replied immediately, "Call the young man Heavenly Smith, and me Chemistry Smith."

Dr. Smith was a man of deep emotions. When the writer suffered a serious accident several years ago his friend stood by his bedside, and with tears in his eyes said, "What a shame! What a dirty shame!" Here was no rebellion at an act of Providence, but the spontaneous expression of sympathy for a stricken comrade.

Dr. Smith's interests extended beyond the classroom and laboratory. He spent his vacations, and his meager savings, traveling to far away places not listed by tourist agencies. His interest in religion was as sincere and vital as that in science. He was faithful in his chapel attendance, and when called on to lead he frequently recited whole chapters of the Scriptures. He did so not to display his disciplined memory, but to impress students with the importance of Holy Writ. He served three years as president of the Baptist General Association of Michigan, a rare distinction for a layman and a scientist.

Lemuel F. Smith

Such a dynamic life evolves its own philosophy of life. On his desk Dr. Smith left these gems of distilled wisdom: "Life is to be judged qualitatively, not quantitatively"; "Everything matters, but nothing too much"; "Nothing is ever finally settled"; "The evening of life gathers into fruits all the morning dews and noonday heat."

[1] KNAPP, R. H., AND H. B. GOODRICH, *Science*, **113**, 543 (1951).

NUCLEAR CHEMISTRY IN THE CURRICULA OF AMERICAN COLLEGES AND UNIVERSITIES[1]

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In a recent issue of *Chemical and Engineering News* appeared the statement, "Five per cent of the total scientific and engineering population of the United States is engaged in atomic energy work." This illustrates the rapidity of the rise of an industry which did not exist ten years ago and forecasts the employment of a significant fraction of our present students. Even a casual examination of scientific periodicals of recent years will also demonstrate a corresponding research effort in pure and applied nuclear chemistry. It is therefore appropriate that we should consider the extent to which our students are being trained or are being offered training in the basic subject matter of this field.

It is not to be expected that collegiate training should have for its purpose the training of specialists in this or other limited fields of science and technology, It is proper, however, to include instruction in the basic concepts of an important field of chemical science at the introductory level. This will serve as a basis for further training of those who enter these fields and will give important breadth to the training of others. Thus, the elementary principles of nuclear phenomena and related subjects have a useful place in the chemistry curriculum at all levels; as a minor part of any general chemistry course, as a somewhat more substantial part of intermediate courses dealing with the structure of matter (such as physical and inorganic chemistry), and finally as a specialty course, offered but not required at the senior or graduate level. Needless to say, both lecture and laboratory instruction in this subject, as in any other, is important.

For the purposes of this discussion we will use the term "nuclear chemistry" to refer to all subject matter dealing with nuclear structure and phenomena which may appear in courses offered by departments of chemistry. Small amounts of such material, more properly described as nuclear physics, have appeared in elementary chemistry texts for some years. Recent editions have shown increased emphasis on the subject although the main part of the material is still concerned with physical rather than chemical phenomena. With respect to advanced lecture instruction, adequate textual material has recently appeared in the form of texts entitled "Radiochemistry," "Nuclear Chemistry," or "Nuclear Physics" (*1*). These may be used as basis for lecture courses completely devoted to nuclear chemistry (*2*) or for the introduction of selected portions of this material into more general courses.

A number of institutions have developed laboratory procedures for advanced instruction in nuclear chemistry (*2*) and at least one published version of such material is available (*3*). Some recent editions of elementary chemistry laboratory manuals contain procedures for nuclear chemistry experiments, but most of these are not very satisfactory.

THE QUESTIONNAIRE

In an attempt to assess the amount of training in nuclear chemistry now given or available to undergraduate chemistry majors, and also some of the factors influencing changes in these programs, we have prepared and distributed a questionnaire on the subject to more than 400 departments of chemistry throughout this country, namely, all those at schools listed as having an enrollment greater than 1000 in the 1949 "World Almanac." About 200 replies were received and these may represent a selection in favor of those interested in the subject.

Table 1 summarizes the general averages of hours of instruction in nuclear chemistry in various courses. From these data it is immediately clear that in the basic courses lecture instruction in nuclear chemistry has spread much more widely than laboratory instruction. However, where laboratory instruction is being undertaken, it appears to be receiving the proper proportionate emphasis. Much the same may be said about advanced instruction although it must be remembered that not all students will take the advanced courses even when they are available.

TABLE 1
General Survey of Instruction in Nuclear Chemistry
(Total number of replies, 192)

		Schools teaching material, %	*Average where taught, hr.*	*Overall average, hr.*
Gen. chem.	*a.* lect.	93	3.9	3.6
	b. lab.	8	3.3	0.3
Phys. chem.	*a.* lect.	76	4.7	3.6
	b. lab.	14	6.2	0.9
Adv. chem.[a]	*a.* lect.	51	15.8	8.1
	b. lab.	11	30	3.5

[a] Includes advanced physical, advanced inorganic, nuclear chemistry, etc.

We have attempted to construct, from the data

[1] Presented at the XIIth International Congress of Pure and Applied Chemistry, New York, September, 1951. The work was supported in part by the Atomic Energy Commission under Contract At(11-1)-38.

supplied by the questionnaire, a more detailed picture of the amount of instruction in nuclear chemistry received by undergraduate chemistry majors. In compiling this information we have considered not only the amount of instruction per school, but also the amount per student, as measured by the size of the June, 1950, graduating class of chemistry majors. Although this number may be expected to fluctuate considerably from year to year at a given school, in a large group of schools it is considered the best measure of capacity in chemical education, and more reliable in this respect than total enrollment. Figure 1 gives the information on lecture instruction in the basic courses, which all chemistry graduates are presumed to have taken. In the same figure the total instruction in basic plus advanced courses is given in terms of the number of undergraduates enrolled in the advanced course (usually a much smaller number than the total number of graduates). Figure 2 represents similar information with respect to laboratory instruction. From these figures, therefore, the per cent of students receiving and schools offering a specified amount of instruction may be obtained. In the case of lecture work, the figure indicates that half of our graduates are receiving six to eight or more hours of instruction in nuclear chemistry in the basic courses, but in laboratory work 76 per cent of the students receive no instruction at all. This disparity between lecture and laboratory work is equally evident in the smaller group of students who study nuclear chemistry in advanced courses. On the other hand the total numbers of lecture and laboratory hours devoted to all topics in the basic courses are normally about equal, and therefore we may conclude that introduction of nuclear chemistry into laboratory work has not kept pace with its introduction into lecture work. This is not unexpected in a relatively new subject, but indicates a direction for progress.

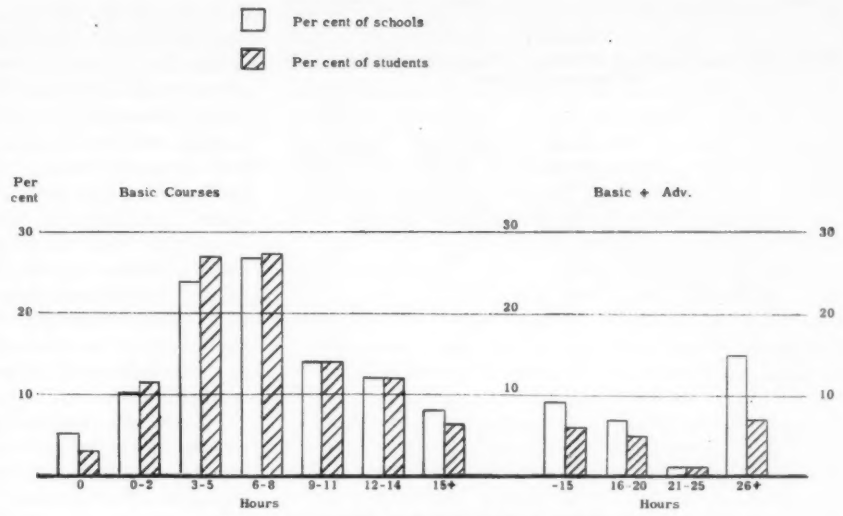

Figure 1. Lecture Training in Nuclear Chemistry

Where the per cent of students receiving a specified amount of instruction significantly exceeds the per cent of schools offering such instruction, as in the case of basic laboratory work (Figure 2) we may conclude that this instruction is more prevalent in schools graduating large classes of chemists. Such a difference is not noticeable in the corresponding lecture work. Where the per cent of students receiving a specified amount of instruction is significantly less than the per cent of schools offering such instruction, as in the case of advanced lecture and laboratory work, we may conclude that a small fraction of the undergraduates have chosen the courses concerned, since these will normally be electives.

Since the information reported in Figures 1 and 2 with respect to advanced courses represents the total student training in nuclear chemistry in basic plus advanced courses, a separate analysis of the advanced courses is given in Table 2.

The information given in Table 2 indicates that, with respect to advanced lecture work, a large percentage of students is offered a small amount of instruction, probably as part of an advanced course in physical or organic chemistry, and that almost two-thirds accept this offering. Courses containing substantial amounts of nuclear chemistry are much less available and also are chosen by only about one-third of the students. Very little advanced laboratory work is available to or accepted by undergraduates

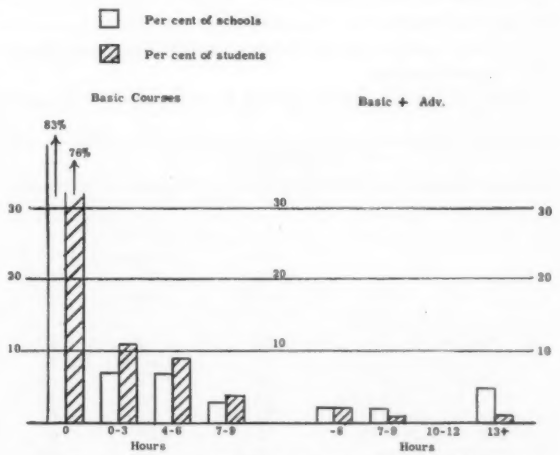

Figure 2. Laboratory Training in Nuclear Chemistry

TABLE 2
Potential and Actual Nuclear Chemistry Training in Advanced Courses

	I *% of schools offering this instruction*	II *% of graduates offered this instruction*	III *% of graduates receiving this instruction*
Lecture hours			
1–10	26	30	18
11–20	11	9	4
21–30	6	7	2
31+	8	10	3
Total	51	56	27
Lab. hours			
1–10	5	5	3
11–20	2	2	1
21–30	0	0	0
31+	5	8	2
Total	12	15	6

and the proportion of available laboratory work to available lecture work is very low as in the basic courses. With the exception of the substantial advanced laboratory work (31+ hours) no significant effect of school size appears in this table.

In addition to providing information on the present status of instruction in nuclear chemistry the questionnaire requested data which might be used to assess certain factors connected with the amount of instruction given. In Table 3 data are presented which permit examination of the possibility of correlation between nuclear instruction and the size of the June, 1950, chemistry class, the presence of a staff member either experienced or active in the field, and American Chemical Society accreditation. A preliminary examination revealed no evidence of correlation between amount of lecture work in the elementary courses and any of these factors and therefore they are omitted from the table. In addition the advanced lecture and laboratory courses have been divided into two easily recognized categories: those in which nuclear chemistry is a part of a more general course and the nuclear chemistry course proper.

From the information given in column I it appears that only in the categories of extensive nuclear chemistry laboratory work is there an evident dependence on the size of the institution (as measured by the size of the graduating class of chemistry majors). In view of the investment of money, space, effort, etc. involved in such work, this finding is not unexpected. In specialized advanced lecture courses, and in introduction of nuclear work into general chemistry laboratory courses, some superiority of larger schools is apparent. In connection with the presence of experienced or active staff, a similar correlation holds, except that the influence is directed toward physical chemistry laboratory, rather than general chemistry. Schools which have A. C. S. accreditation again appear strong in the extensive advanced laboratory and lecture courses, with smaller influence in physical chemistry laboratory.

TABLE 3
Factors Connected with Amount of Nuclear Instruction

	I *Average chem. grad. class*	II *% with active or exper. staff*	III *% A. C. S. accredited*
Schools teaching nuclear in: Gen. chem. lab.	23	73	60
Phys. chem. lab.	21	89	74
Adv. lect. (1–20 hr.)	20	72	58
Adv. lect. (>25 hr.)	24	88	92
Adv. lab. (1–15 hr.)	21	82	64
Adv. lab. (>40 hr.)	31	100	89
All schools reporting	18.5	65	48

Finally, the questionnaire inquired what increases in nuclear instruction were contemplated and what factors seemed most important in discouraging such changes. The answers to the first of these questions are summarized in Table 4 and it appears that the increases will be much more evenly divided between lecture and laboratory work than in present instruction and that these increases, where made, will be quite substantial. The contemplated increases are mainly in advanced courses which not all students will take, and in only one case, advanced laboratory work, will the increase produce a substantial change in the overall average (compare figures in parentheses taken from Table 1).

TABLE 4
Planned Increases in Nuclear Instruction

	Schools increasing, %	*Average increase (where planned), hr.*	*Average increase (over 192), hr.*
Basic courses *a.* lect.	4	2	0.1 (7.2)
b. lab.	7	3	0.2 (1.2)
Adv. courses *a.* lect.	14	12	1.6 (8.1)
b. lab.	10	18	1.9 (3.5)

TABLE 5
Discouraging Factors

	Times mentioned	*Average grad. class*
Cost	99	18
Laboratory space	57	21
Experience	45	
Desirability	45	
Curriculum	31	
Hazard	31	

(Only 178 completed this part of the questionnaire)

In Table 5 are listed some of the discouraging factors offered to the respondents for selection. They were asked to rank such factors in order of importance and to add to the list any others considered important. The number of first or second rankings is given in each case and it appears that cost (arbitrarily specified to be a minimum of approximately $300 for installation of

quantitative self-contained nuclear laboratory experiments) is most discouraging. For those mentioning this item and the next most important, lack of laboratory space, the average size of the June, 1950, class was computed, in order to test for correlation with school size. Apparently no such correlation exists. Lack of time in the curriculum was frequently mentioned, although not specifically offered in the questionnaire.

SUMMARY

The results obtained from the questionnaire indicate that instruction in nuclear chemistry by lecture is a small but widespread part of the general and physical chemistry courses. Many chemistry students have the opportunity for a few more hours instuction by lecture in some advanced course and as many as 10 per cent of all students are offered an extensive lecture course in nuclear chemistry. Apparently the instructors find the basic texts now available fairly adequate, since no strong correlation with staff experience or activity in the field appears except for the specialized courses where such factors would be expected to influence any specialized course. One may question the efficiency of dividing the elementary lecture instruction into small units associated with several courses as is commonly done, but the total amount of time, if properly managed, should suffice.

Laboratory instruction, especially in the basic courses, has not kept pace with lecture instruction. Very few of our graduating chemists have had even a single period of laboratory instruction in nuclear chemistry. However, in the few schools where laboratory instruction is offered, it apparently assumes the proper quantitative relation to lecture work. These schools are, in the case of nuclear instruction in the general chemistry laboratory, larger than average, and, in the case of physical chemistry laboratory, predominantly those with staff members experienced or active in the field. With respect to advanced and especially the extensive courses in nuclear chemistry, laboratory instruction is relatively well established at the larger schools and those where experienced or active staff members are available.

The increases planned in lecture and laboratory instruction in nuclear chemistry appear to be well balanced, although they will not be widespread. These increases will be most significant in the advanced courses. It is encouraging to note that the small increases in the basic course work are more in laboratory work than in lecture. However, the advanced laboratory work will improve more than will elementary laboratory work.

It seems clear that the most obvious shortcoming of our present teaching program with respect to nuclear chemistry is the lack of elementary laboratory work. Regardless of the absolute importance assigned to the subject, it is obvious that the elementary course work has a disproportionately small amount of laboratory instruction. With regard to the discouraging factors offered by the questionnaire, we fail to understand the apparent importance of the two items "lack of conviction of desirability" and "lack of time in curriculum," since these should not prevent a proper balance of laboratory and lecture instruction. We believe that the discouraging factors "possible hazard" and perhaps also "lack of laboratory space" may have been overrated by the respondents, due to the very real factor of "lack of experience." This last item certainly discourages routine adoption of a new and strange technique, and it must be aggravated by lack of suitable textual material for elementary experiments. Given the latter, a considerable amount of lack of experience can be tolerated in elementary instruction. "Cost" appears to be uppermost in many minds, but this is also connected with the preceding items. Surely the investment of as much as $300 in permanent equipment for a physical chemistry experiment is not uncommon and could certainly be justified if the associated textual material were easily available. Such a cost is, to be sure, rather low in comparison to advertised prices of nuclear equipment, but it should be readily attainable if the work is widely adopted.

In conclusion, we believe that the most urgent need for progress in nuclear chemistry instruction is in regard to laboratory work in the basic courses. Laboratory experiments now being used in advanced instruction (*2*, *3*) must be judiciously adapted for elementary instruction on a larger scale. Some attempts to meet this need have recently been described (*4*, *5*). However, much room for improvement and invention of elementary procedures remains. Suitable textual material is needed and reasonably priced equipment must be made available (*6*) before any great improvement can be expected. We urge that the Atomic Energy Commission, together with the appropriate commercial and academic institutions, act on this need to improve the training of our chemists in an important field.

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WILHELM KLEMM

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THE internationally known inorganic chemist, Wilhelm Karl Klemm, was born on January 5, 1896, at Guhrau in eastern Germany. He entered the nearby University of Breslau in 1919 and received the Ph.D. degree in 1923. His dissertation, which dealt with uric acid, was prepared under the guidance of Heinrich Biltz (1865–1943), who sent the promising young man to the Technische Hochschule at Hannover for further training under his brother Wilhelm Biltz (1877–1943).[1] Dr. Klemm was assistant until 1927 when he became Privatdozent and then Associate Professor in 1929. He headed the inorganic laboratory at Danzig from 1933 until 1945, when the conquest of the city by the Soviets made flight necessary. He was at the University of Kiel in a like capacity, from 1947 to 1951, when he came to his present post, Director of the Institut für anorganische Chemie at the University of Münster.

Professor Klemm's researches have been many and varied. His first studies were in collaboration with W. Biltz and dealt with the conductivity of fused salts. They found that the high-melting halides are good conductors, in contrast to the low-melting ones, which conduct poorly; clear relationships to the periodic system of the elements were brought to light. The investigations were extended to the study of the molvolumes and the coefficients of expansion of the fused halides, and the data obtained likewise gave insights into the molecular states obtaining in the melts. Furthermore, the expansion coefficients of the solid halides were measured and again systematic relationships were uncovered. As a supplement to these studies, the halogen and chalcogenic compounds of several uncommon elements (indium, gallium, rare earths, germanium, rhenium) were investigated in detail; the heats of formation and the reactivity toward ammonia were determined.

Study of the compounds of the rare earths led to the discovery of bivalent ytterbium. Moreover, a new method for classifying the rare earths was evolved, a system that has now been quite generally adopted. The decisive factor was recognized to be the stability not only of the completely filled but also the "half-filled" electron configuration, such as is found in Gd^{+3}. Subsequently, this systematization was applied to the transition elements, and the special properties of Mn^{++} compounds were explained. The new classification of the rare earths, which at the start held only for the ions, was found to be valid also for the metals themselves. The latter were prepared, with the collaboration of H. Bommer, by means of the reaction: $MeCl_3 + 3K = 3KCl + Me$, and studied roentgenographically and magnetically. The atomic volume curves of the rare earths, which could then be constructed for the first time, show distinct maxima in the case of ytterbium and europium and flat minima for cerium, praseodymium, and terbium. Like gadolinium, practically all the other yttria earths are ferromagnetic at low temperatures.

Magnetic measurements, which had been employed in the above investigations, were used, from 1930 on, to a much greater degree for the solution of chemical problems. In addition to other compounds of the rare earths, numerous complex compounds were measured, particularly simple compounds of the transition elements, including those with semimetallic character. The importance of atomic bonding between the cations (antiferromagnetism) was recognized independently of other workers, and applied liberally to account for the experimental findings. A classification system was developed also for the diamagnetism of ions.

Studies of alkali metal compounds led to, among other things, the discovery that the O_2^- ion occurs not only in dioxides but also in the oxides Rb_2O_3 and Cs_2O_3, whose structures were thus elucidated.

In the field of intermetallic systems, studies were made of the behavior of the alkali and alkaline earth metals toward each other, and toward magnesium, and many systems with semimetals were also investigated. For the transition from the true metals to the semimetals it was proposed that the following categories be adopted: true metals, metametals, semimetals, nonmetals. The decrease in volume when gallium, silicon, germanium, and bismuth melt was explained. The volume-temperature curve of tellurium was found to have a minimum.

In another important series of researches, definite evidence was obtained for the existence of gaseous compounds of univalent aluminum with all halogens, sulfur, and selenium.

Numerous studies were devoted to fluorine compounds. For example, measurements were made of the refractive power of SF_6 etc.; CeF_4 and the fluorides of gallium and indium were prepared. Complex fluorides which have abnormal valence, such as $K_3[Cu(III)F_6]$, $K_2[Ni(IV)F_6]$, $K_2[Cr(IV)F_6]$, etc., were produced by the action of fluorine on various appropriate compounds or mixtures.

In addition to his many papers, Professor Klemm has authored or collaborated on a number of books. These include: "Magnetochemie" (1937), which was reprinted at Ann Arbor in 1944 as Volume I of "Physik und Chemie und ihre Anwendungen"; and "Chemische Binding als electrostatische Erscheinung"

[1] KLEMM, W., *Z. anorg. allgem. Chemie*, **231**, 3–14 (1937).

(1931) which is his translation of A. E. van Arkel and J. H. de Boer's "Chemische Binding als elektrostatische Verschynsel." He collaborated with W. Biltz on the latter's "Raumchemie der festen Stoffe," which likewise was reprinted at Ann Arbor in 1944. For student use he published an "Introduction to Inorganic Chemistry" (5th edition, 1952). Together with Werner Fischer, he issued several revisions of the Biltz "Experimentelle Einführung in die anorganische Chemie," which has now gone through more than 40 editions.

Professor Klemm is one of the three chief editors of the *Zeitschrift für anorganische und allgemeine Chemie.* Furthermore, he edited the six volumes of the inorganic section of the "Fiat Review," and also "The Review of Analytical Chemistry in Germany 1936–1946," an important record of German chemistry during the World War II era.

His honors include the Justus von Liebig Medal of the Gesellschaft Deutscher Chemiker (1951), whose presidency he holds for 1952 and 1953.

COULOMB'S LAW AND THE QUANTITATIVE INTERPRETATION OF REACTION RATES[1]

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THE considerations discussed in this paper apply only when the solvent has no specific effect on the reaction between chemical reactants and exerts its influence as an electrostatic effect through its dielectric constant.

In a former article (*1*) Coulomb's law and the qualitative interpretation of reaction rates was discussed. The Coulomb approach is applicable to the quantitative interpretation of reaction rates as influenced by the dielectric constant of the medium. The application in a quantitative manner of Coulomb's law is easiest seen when the force equation is used to obtain the difference in the energy of activation for a reaction which takes place at two different constant dielectric constants of the same solvent pair. Again, a solvent pair (*1*) must be used in order to obtain the desired dielectric constants in which the reaction is to take place. Thus, one constant dielectric constant chosen might be that of water at 40°C., namely, 73.1. As can be found from Åkerlöf's (*2*) data, a solution of 12.5 per cent by weight of ethanol in water has a dielectric constant of 73.1 at 20°C. If the specific rate constant of a reaction is determined at 40° in pure water and 20° in 12.5 per cent ethanol-water solvent using the same concentrations of reactants, the dielectric constants of the two media will be the same and, therefore, the temperature coefficient obtained will be independent of electrostatic effects arising from changing dielectric constant. This type of temperature coefficient will arise from more purely thermal and nonelectrostatic effects. Again let us run the same reaction at the same concentrations of reactants in 10 per cent by weight ethanol in water solvent at 40° and in 21.3 per cent by weight ethanol in water solvent at 20°. From Åkerlöf's data the dielectric constant of these two solvents will be the same and will have the value of 67.9, so that in this case also the temperature coefficient will be independent of electrostatic effects arising from changing dielectric constant.

Now suppose that the energy of activation is calculated from the rate constants at the two temperatures when the dielectric constant is constant and equals 73.1 and that a similar calculation is made from the data at constant dielectric constant 67.9. The two energies of activation will not be equal but will differ quantitatively by an amount equivalent to the difference in Coulombic energy necessary to bring the two reactant particles to within the distance r of each other; where r is the distance separating the particles necessary for reaction to take place. From this difference in activation energies, a quantitative difference in the specific rate constants at the two dielectric constants can be obtained.

Two points need clarifying at this time. One is an explanation of how energy of activation is calculated from rate constants and the other is how the quanti-

[1] This work was done on a contract made by the Institute of Science and Technology of the University of Arkansas with the Office of Naval Research.

tative difference in energies of activation, *i. e.*, Coulombic energies of activation, can be found from Coulomb's law.

Energy of activation is generally calculated from specific rate constants determined at different constant temperatures and using the Arrhenius equation. This equation is

$$\log k' = \log Z - \frac{\Delta E}{2.303RT} \tag{1}$$

when decadic logarithms are used. In this equation k' is the specific velocity constant of the reaction, Z is the Arrhenius frequency factor, ΔE is the energy of activation per mol, R is the molar gas constant in calories if ΔE is to be calculated in calories, and T is the absolute temperature. The energy of activation ΔE is the energy per mol that reactants must receive if they are to form products.

From equation (1) two ways suggest themselves for the determination of ΔE. The better way is to find the specific rate constant for the reaction at several values of T. The logarithm of k' is then plotted against $1/T$ and a straight line is obtained if Z and ΔE, are independent of T, which they generally are for short ranges of T. The slope of this line, which is always negative (*3*), is set equal to $-\Delta E/2.303R$, and ΔE is calculated from this equality.

A less accurate way of calculating ΔE is to choose two values of $\log k'$ at two values of T and substitute in the formula

$$\Delta E = \frac{2.303T_2T_1\ \Delta \log k'}{\Delta T} \tag{2}$$

obtained from equation (1) by subtracting $\log k'$ at T_2 from $\log k'$ at T_1 and the term containing $1/T_2$ from that containing $1/T_1$ assuming $\log Z$ constant (independent of T).

Coulombic energies of activation, *i. e.*, the change in energy of activation necessary to bring the two reactant particles to within reacting distance at two different constant dielectric constants, preferably of the same solvent pair to obviate even any minor specific solvent effects, can be calculated for ionic reactants in the following manner. Coulomb's force equation applied to ions in a solvent of dielectric constant D is

$$\mathrm{F} = \frac{z_1\,\epsilon\ z_2\,\epsilon}{Dr^2} \tag{3}$$

where F is the force between two ions at distance r apart in a dielectric media of dielectric constant D, if the valences of the ions are z_1 and z_2. ϵ is the electronic charge, and therefore $z_1\,\epsilon$ and $z_2\,\epsilon$ are the ionic charges.

The Coulombic energy is given by the expression

$$E = -\int_{\infty}^{r} Fdr = \int_{r}^{\infty} Fdr = \frac{z_1\epsilon\ z_2\epsilon}{D}\int_{r}^{\infty}\frac{dr}{r^2} \tag{4}$$

Therefore

$$E = \frac{z_1\epsilon\ z_2\epsilon}{D}\left[-\frac{1}{r}\right]_{r}^{\infty} = \frac{z_1\epsilon\ z_2\epsilon}{Dr} \tag{5}$$

and the difference of Coulombic energy in changing from dielectric constant D_1 to dielectric constant D_2, *i. e.*, the molar Coulombic energy of activation, is

$$\Delta E_c = \frac{N\ z_1z_2\epsilon^2}{r}\left(\frac{1}{D_2} - \frac{1}{D_1}\right) \cdot \tag{6}$$

Equation (6) and corresponding ones for other electrical charge types of reactants have been given by Amis (*4*, *5*) and verified from data.

To state the influence of Coulombic energy of activation on specific reaction rates quantitatively consider a reaction at dielectric constant D_2 and let its energy requirements be represented by ΔE_{D_2}. Let the energy of activation for the same reaction at dielectric constant D_1 be represented by ΔE_{D_1}, then

$$\Delta E_{D_2} = \Delta E_{D_1} + \Delta E_c \tag{7}$$

Now

$$k'_{D_2} = Ze - \frac{\Delta E_{D_2}}{RT} \tag{8}$$

and

$$\log k'_{D_2} = \log Z - \frac{\Delta E_{D_2}}{2.303RT}$$

$$= \log Z - \frac{\Delta E_{D_1}}{2.303RT} - \frac{\Delta E_c}{2.303RT} \tag{9}$$

but

$$\log k'_{D_1} = \log Z - \frac{\Delta E_{D_1}}{2.303RT} \tag{10}$$

therefore,

$$\log k'_{D_2} = \log k'_{D_1} - \frac{\Delta E_c}{2.303RT} \tag{11}$$

and substituting the value of ΔE_c from equation (6) in equation (11) one obtains

$$\log k'_{D2} = \log k'_{D_1} - \frac{N\ z_1z_2\epsilon^2}{2.303RTr}\left(\frac{1}{D_2} - \frac{1}{D_1}\right) \tag{12}$$

If a reference dielectric constant D_2 is taken as infinite in magnitude so that all electrostatic forces between reactants vanish and if D_1 is taken as any dielectric constant D, we have from equation (12), remembering that N/R is equal to $1/k$

$$\log k'_{\infty} = \log k'_D + \frac{z_1z_2\epsilon^2}{2.303kTr\ D} \tag{13}$$

or

$$\log k'_D = \log k'_{\infty} - \frac{z_1z_2\epsilon^2}{2.303kTr\ D} \tag{14}$$

k'_{∞} which is at zero ionic strength, is independent of D since all electrostatic forces between reactant ions vanish at infinite dielectric constant. Hence as D increases equation (14) predicts an increase in k'_D for reactant ions of like charge sign and a decrease in k'_D for reactant ions of unlike charge sign. As D decreases k'_D will decrease for ions of like sign and vice versa for ions of unlike sign.

A plot of log k'_D at zero ionic strength versus $1/D$ should give a straight line in the region of dielectric constant D where theory is being obeyed. This line should have a negative slope for ions of like charge sign and a positive slope for ions of unlike charge. Scatchard (*5*) derived an identical expression for the influence of the dielectric constant of the medium on reaction velocity constants from the direct calculation of the concentration of an intermediate complex from equations given by Debye and Hückel. Here we have arrived at the expression from the standpoint of the simple expression for the Coulombic force between ionic reactants.

Scatchard's approach was like that of Christiansen who arrived at an expression for the influence of ionic strength upon reaction rates. The combined effect of ionic strength and dielectric constant upon reaction rates is given by the Brønsted-Christiansen-Scatchard equation (*4b*). The equation for dilute solutions is

$$\log k' = \log k'_{\substack{D = \infty \\ \kappa = 0}} - \frac{z_A z_B \epsilon^2}{DkTr} + \frac{z_A z_B \epsilon^2}{DkT}\kappa \qquad (15)$$

In this equation k' is the specific velocity constant of the reaction at finite dielectric constant and ionic strength,

$$k'_{\substack{D = \infty \\ \kappa = 0}}$$

is the specific velocity constant extrapolated to a dielectric constant of infinity and a kappa (κ) of zero, κ is the Debye Kappa and is proportional to the square root of the ionic strength, μ, and the other quantities have been defined. This equation has been verified many times with respect to both the dielectric constant and ionic strength dependence of reaction rates (*4b*). Thus equation (14) has also been verified. The dielectric constant dependence was verified as discussed in relation to equation (14) by plotting log k' at zero ionic strength versus $1/D$. The ionic strength dependence was checked by plotting log k' at constant finite dielectric constant versus the square root of ionic strength and proving in dilute solutions that the straight line obtained had the direction and magnitude of slope predicted by equation (15) when κ was substituted for in terms of ionic strength (See ref. (*4b*), Chap. 4).

Thus ion-ion types of reactions have been discussed from both qualitative and quantitative standpoints. Quantitatively the Coulomb's law approach allows only an interpretation of the dielectric constant effect upon reaction rates. This approach is simple in that it does not involve an intermediate complex but deals only with Coulomb forces and energies between the reactant ions themselves. The method is also applicable to combinations of ions and electrically unsymmetrical molecules and to combinations of different electrically unsymmetrical molecules. These other combinations may be dealt with in a later paper.

LECTURE ILLUSTRATION

. Since time will not permit actual kinetic runs to be made and specific velocity constants determined during a lecture period, a quantitative demonstration of the electrostatics of an ion-ion reaction shall be in the form of a problem using the kinetic data of Svirbely and Schramm (*6*) and of Lander and Svirbely (*7*) for the reaction at 30°C. in water and glycol-water solvents of the ammonium ion and cyanate ion to give urea. The over-all reaction is

$$NH_4^+ + OCN^- \rightarrow (NH_2)_2CO$$

The kinetic data are given in the table.

Kinetic Data for the Reaction at 30.0°C. in Water and Glycol-Water Solvents of the Ammonium Ion and Cyanate Ion to Give Urea

Wt.% glycol	*Dielectric constant D*	$k'_{\kappa = 0}$
0.00	76.73	0.00620
44.25	63.5	0.0113
54.20	60.0	0.0140
68.0	55.0	0.0191
76.15	50.0	0.0265
85.75	45.0	0.0378
94.10	40.0	0.0509

Plot values of log $k'_{\kappa = 0}$ as ordinates versus values of $1/D$ as abscissas. Down to $D \cong 50$ ($1/D \cong 0.02$) the line will be straight with a positive slope as predicted by equation (14). The deviation from a straight line below D $\cong$ 50 has been mentioned before (ref. (*4b*), Chap. 4). Determine the slope of the straight portion of the line and, by equating this slope to

$$-\frac{z_1 z_2 \epsilon^2}{2.303kTr}$$

as required by equation (14), prove the value of r to be reasonable, *i. e.*, a molecular dimension. Remember that in this case, where a positive ion is reacting with a negative ion, the z_1z_2 product is negative. The answer is $r = 2.5 \times 10^{-8}$ cm.

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(5) Scatchard, George, *Chem. Revs.*, **10**, 229 (1932).
(6) Svirbely, W. J., and A. Schramm, *J. Am. Chem. Soc.*, **60**, 330 (1938).
(7) Lander, J., and W. J. Svirbely, *ibid.*, **60**, 1613 (1938).

THE VALUE OF THE HISTORY AND PHILOSOPHY OF SCIENCE IN THE TRAINING OF GRADUATE STUDENTS IN CHEMISTRY[1]

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"THEY that know the entire course of the development of science will, as a matter of course, judge more freely and more correctly of the significance of any present scientific movement than they who, limited in their views to the age in which their own lives have been spent, contemplate merely the momentary trend that the course of intellectual events takes at the present moment."[2]

This quotation from the introduction to Ernst Mach's "The Science of Mechanics" really sums up the value of the history of science and the history of its philosophies to the graduate student in any field of science.

If any criticism of this statement is made, it might be that it is not sufficiently comprehensive, for it leaves out aspects of life not directly concerned with the student's scientific work. If it is assumed that the graduate student is not only pursuing a course of study to attain professional competence, but is also losing no opportunity to come to an understanding of the sciences related to his special field, and an understanding of the world he lives in today, then, it seems to me, that he can achieve this understanding by knowing something of how it all came about. This is primarily the history of science, an essential part of the history of civilization. The student might well ask, "But why bring in the philosophy of science?" The answer is simple: "Because it cannot be left out." The history of science is the study of the discovery of facts and of the ideas of their relations. A collection of facts does not constitute a science; the important aspect of facts about the universe around us is the ideas that grow out of thinking about their relations to each other. As soon as man began to think about the relationship of facts he became a philosopher. His interpretation of the facts became his philosophy. And so it happens that whether the historian of science is conscious of having a philosophy of science or not, every word that he writes about the history of science reveals that he has such a philosophy, be it good or bad, modern or old-fashioned.

Perhaps it would be well at this point to define two terms, both hard to define: science and the philosophy of science. Books have been written about both, mostly, by the way, by philosophers, physicists, and mathematicians. So far as I know Wilhelm Ostwald, founder of modern physical chemistry, is the only chemist who wrote extensively of the philosophy of science.[3] What is science? I shall give two definitions. Henri Poincaré, a great mathematician, says, "Science is a system of relations. . . . Objectivity must be found in relations. . . . Eternal objects are really objects, groups of sensations cemented by a constant bond. Science does not teach the true nature of things but the true relations of things."[4] In this definition the essence of science is linked with reality. This is an old problem—the problem of what is real in the world. This is metaphysics, defined by another mathematician-philosopher, Bertrand Russell, as "the attempt to conceive the world as a whole by means of thought."[5]

As a result of the thinking and writing of another mathematician-philosopher, Alfred North Whitehead, another definition of science has come into vogue in the last 15 years: Science is a process that began before recorded history, one that is still going on, and that will continue as long as man is man. This is the aspect of science emphasized by James B. Conant,". . . Science emerges from the other progressive activities of man to the extent that new concepts arise from experiments and observations, and the new concepts in turn lead to further experiments and observations."[6] According to this idea science has a dynamic quality like the universe it proposes to explain; it is itself forever in a condition of flux or change. According to this view an understanding of any science necessarily involves some knowledge of the history of its development.

The definition of the philosophy of science I shall give is adapted from Whitehead's definition of philosophy.[7] The philosophy of science is an attempt to clarify those fundamental beliefs about the nature of the world which finally determine the emphasis of attention of the scientist that lies at the base of his scientific achievements. What bearing has this definition on the work of a graduate student in chemistry? It is all important for

[1] Presented before the Chemistry Journal Club of Tulane University, February 7, 1952.

[2] MACH, ERNST, "The Science of Mechanics," trans. by T. J. MCCORMACK, Open Court Publishing Co., LaSalle, Ill., **1942**, p. 8.

[3] Perhaps one should mention also G. N. Lewis whose excellent book, "The Anatomy of Science," appeared in 1926 (Yale University Press).

[4] POINCARÉ, H., "The Foundations of Science," trans. by G. B. HALSTED, The Science Press, New York, **1921**, p. 349.

[5] RUSSELL, B., "Mysticism and Logic," W. W. Norton, New York, **1929**, p. 1.

[6] CONANT, J. B., "On Understanding Science," Yale University Press, New Haven, Conn., **1947**, p. 24.

[7] WHITEHEAD, A. N., "Adventures of Ideas," Macmillan Co., New York, **1933**, p. 125. "Philosophy is an attempt to clarify those fundamental beliefs which finally determine the emphasis of attention that lies at the base of character."

it will be the deciding factor determining the field of chemistry in which he specializes, the way he approaches his research problem, the kind of work he does, the success that attends his accomplishments after he gets his degree.

It is a rare student in chemistry who begins graduate work knowing much about himself, his talents, his ability, his weaknesses. I might also add the subject itself, for he knows little or nothing of its history and the history of the sciences closely related to it. The courses in chemistry today are full to overflowing with the facts of the chemistry of the present; there is no time to do more than mention some of the theories that once held the attention of the great scientists, whose laboratory investigations and whose thinking are responsible for the remarkable achievements of our time. Often even the names of the founders of chemistry are omitted to make place for the names of discoverers of recent additions to the development of the science. This is particularly true of organic chemistry; the graduate student in this field who has a fairly retentive memory can rattle off the names of men of only moderate stature who happened to discover a new method of synthesis or reduction, and be totally unaware of the contributions or methods of attack of really great experimentalists like Auguste Laurent, Adolph Wurtz, and Emil Fischer. And, of course, such a student would be at a loss to understand even the point of mentioning an hypothesis of one of the intellectual giants of the seventeenth century. To most students all important ideas are derived from the science of very recent years.

Reading and study in the history of science and particularly in the history of chemistry would give the student the background essential for the evaluation of the theories of today. He would learn also much about his own interests, even his own capacity for original thinking and his own shortcomings. It is not necessary to take a course or courses in the field. This, however, is an easy way to get an introduction to what can be read or studied and where to find the information. The problem of what to read to best advantage is important. For the beginner I would suggest biographies, long and short, interestingly written. These should be followed by a detailed study of the work of one particular scientist with the reading of some of his original papers. There are a few excellent books that treat of the development of a theory. One of these is "Avogadro and Dalton" by Andrew N. Meldrum. The particular man or subject chosen, however, really makes little difference. The important thing is that the student lose himself in the development of the thinking and the achievements of a master or masters in the field. In this study certain questions are likely to occur to the student. How did the scientist happen to become absorbed in solving a particular problem? Often this question cannot be answered especially if the man worked in the seventeenth century or earlier, and left no autobiography and no letters that have been preserved. But beginning with the late eighteenth century the origins of many important discoveries have been recorded by the discoverers themselves. Another question: When the problem presented itself what associative idea from his own field or from another suggested to the investigator the method of attack? Another: Was there an element of chance, that is accident, in the solution of the problem? There often is. Pasteur's work on the crystals of the sodium ammonium salt of racemic acid was the result of his association for several months with Auguste Laurent in Balard's laboratory in the Normal School where Pasteur was an assistant.[8] Laurent, an expert crystallographer, who had also worked with optically active alkaloids, not only taught Pasteur how to examine crystals and how to use the polariscope expertly, but he aroused in Pasteur an interest in the problem of the relation of crystalline form to the structure of the molecules.

The view that the history of science is an important aspect of the history or civilization is fairly recent, but even more recent is the attention the subject has been receiving in educational programs of many liberal arts colleges. It seems no exaggeration to remark that the day is not far off when courses in the subject will be offered in most colleges and universities in this country. The delay in organizing such courses is not due to lack of appreciation of their value, but to lack of instructors with some training in several fields of science, and also, more important, with great interest in teaching the subject. Up to now only a few universities in this country—one of them is Harvard—has offered courses in the history of science in certain periods in the development of our western civilization. It is necessary not to confuse a course on the development of science and technology in a particular period, as, for example, in the twelfth and thirteenth centuries, with a course in the history of a particular science. From the time of the Greeks to the present, the development of certain phases of science has received the attention of some few investigators. If this were not true there would today be no record of the medicine of Hippocrates, none of the mathematics of the Pythagoreans, nor of the astronomy of Hipparchus. The writings of these men are lost. What we know of their work we owe to accounts of later investigators. Interest in the history of special sciences was characteristic of most of the great figures of the seventeenth century: William Gilbert in "De Magnete" (1600) includes a history of the lodestone from earliest times, also a history of iron ores and their treatment for the manufacture of iron, also a history of substances like amber which attract light bodies when rubbed; Johannes Kepler in "Dioptrice" (1611) summarizes all experimental work on light from the time of Euclid; Robert Hooke in papers on the telescope and microscope gives lengthy histories of the invention and improvements of these instruments. Little or nothing, however, was written on the history of science as a whole until the second half of the nineteenth century.

Strange as it may seem to scientists, the man whom George Sarton, one of the pioneers in the history of

[8] DUBOS, R. J., "Louis Pasteur," Little Brown, Boston, 1950, p. 90.

science in this country, considers the founder of the history of science was neither a scientist nor a historian but a philosopher with leanings toward sociology. Auguste Comte published his famous "Cours de philosophie positive" from 1830 to 1842. In it he calls attention to three ideas (1) that he could not have accomplished his synthesis had he not depended on the history of science; (2) that in order to understand the history of mankind it is necessary to study the evolution of different sciences; (3) that it is not enough to study the history of one or many particular sciences; one must consider the history of science as a whole. "As early as 1832 Comte made an application to the minister Guizot for the creation of a chair devoted to the general history of the sciences. It was sixty years before this wish of his was granted; and the course entrusted to Pierre Lafitte was inaugurated at the Collège de France in 1892, thirty-five years after Comte's death."[9]

There is no doubt that Comte's writings focused attention on the importance of the subject but the idea must have been in the air, as it were, for there were others who devoted their energies to it. Among these I shall mention two: Justus Liebig and Ernst Mach. Liebig, you recall, became weary of teaching students the art of laboratory investigation in chemistry in the last few years that he taught at Giessen, and when he accepted the call to Munich in 1853, he stipulated that he would take no part in the training of students. Sometime after 1860 Liebig became interested in the theory of knowledge, "by which he wished to render account to himself of that business of discovery in which he had been so long practically engaged. Evidently in order to procure for himself a point of view founded upon history, he had then studied Bacon's writings, who had been designated by professional philosophers in general as the intellectual founder of the age of natural science upon which we entered some centuries ago. The result was a great surprise, for Liebig could find no trace whatever in the teachings of this legal dilettante of the method that was so well known to Liebig from his own practice."[10] The paper which he gave on the subject before the Academy of Sciences in Munich stirred up a "nest of learned wasps." The debates which followed forced Leibig to investigate the character of natural philosophy, as indicated by his own experiences. He came to the conclusion that the "progress of the human race is in reality determined by its inventions which determine its civilization, and by empiric conceptions acquired by the investigation of nature."[11]

Ernst Mach's important contribution to the history of physics and to the philosophical basis of the concepts, now accepted without question by most students, grew out of his teaching of physics and also out of his study of the operations of the mind. His most important work, "The Science of Mechanics, a Critical and Historical Account of Its Development," first appeared in 1883, but in the preface to this first edition Mach states that the book is an elucidation of the ideas presented in a treatise published by him in 1872. He mentions others who wrote on the same theme, among them Kirchhoff and Helmholtz. His interest, he says, was stimulated by reading such things as the introductory chapters of "Analytical Mechanics" by Lagrange. Unlike Liebig, however, Mach must have had a predilection, as Ostwald had, for the history and philosophy of science, for in 1872 he was in his early thirties. An excellent account of the place of Mach in the story of the development of the philosophy of science can be found in Philipp Franck's "Modern Science and Its Philosophy."

Since the middle of the nineteenth century, then, there has been developed, slowly at first, mainly in the universities of continental Europe, this interest in the philosophy of science, especially physics and related sciences. In England and the United States the movement took hold some years later. Among the English should be mentioned William K. Clifford whose lectures on the subject began about 1870,[12] and Karl Pearson, whose "Grammar of Science" appeared in 1892. In the United States among the first were J. B. Stallo whose "The Concepts and Theories of Modern Physics" was published in 1881 and Charles S. Pierce who began publishing papers about this same time.[13]

One does not have to delve deeply into the philosophy of science before one comes to the conviction that some knowledge of its history is essential. Every one of these early writers on the methodology and philosophy of science was conversant with its history. These two aspects of science are almost inseparable. Many who have given thought to the subject advocate a course in the history of science as a prerequisite to study in the history of its philosophies. The two are really one subject. The progress that was made in the development of science from the late sixteenth century to the close of the remarkable seventeenth century, which marks the beginning of modern science, was determined by concepts of matter and energy, space, and time, that arose from an interpretation of the relations of facts derived from observation and experiment. This is nothing more or less than the philosophy of science of that period. The importance of the information about the history of science, that the student brings to an attack on a research problem, has frequently been emphasized by great chemists, especially teachers of chemistry. Willstätter con-

[9] SARTON, G., "Life of Science," Henry Schumann, New York, 1948, p. 30.

[10] OSTWALD, W., "Grosse Männer," Akademische Verlagsgesellschaft, Leipzig, 1910, p. 209.

[11] *Ibid.*, p. 210. For detailed information about Liebig's ideas see the English translation of the address on this subject given before the Royal Academy of Sciences in Munich, July 25, 1866, in "The Culture Demanded by Modern Life" by E. L. YOUMANS, D. Appleton and Co., New York, 1867, p. 347.

[12] CLIFFORD, W. K., "The Common Sense of the Exact Sciences," A. A. Knopf, New York, 1946.

[13] "Charles S. Pierce," ed. by MORRIS R. COHEN, Harcourt, Brace and Co., New York, 1923.

sidered "the teaching and study of the historical development of science indispensable."[14] Today in this country some of the professors directing the study of students are urging that a course in the history of the natural sciences be required. Ernst A. Hauser writes, "To accomplish this broad outlook [he has quoted from Ostwalds' introduction to "The Foundations of Natural Philosophy"] courses in the history of science and technology should be made compulsory subjects for every student in liberal arts and technological colleges and universities."[15]

If the history of science is essential to the training of a graduate student is this also true of the philosophy of science? Some hold that it is the more important of the two. Anthony Standen has recently published[16] some thought-provoking comments on the teaching of chemistry. He points out that only in rare instances does a student learn anything of the method of science even though experimental work is required. The approach to experimental work is mainly factual and very little of reasoning is associated with the doing of experiments. The emphasis in all courses is exclusively on facts, facts that he claims must be known if the graduate is to earn a living in the profession of chemistry. Furthermore he asks: will the student need to know much of the value of an hypothesis and the steps involved in arriving at a satisfactory hypothesis? The answer is no, unless the chemist becomes a director of research. This attitude that facts are all important to the everyday run of professional chemist may have been sufficient some 15 years ago, but the invention of the atomic bomb and the beginning of our difficulty with Russia has forced all scientists, including the chemists, to consider the philosophical foundations of science and their relation to our civilization.

May I remind all students that modern science is one of the products of freedom of thought. This is not a new idea; Liebig emphasized it in 1866: "Freedom, that is, the absence of all restrictions which can prevent men from using to their advantage the powers which God has given them, is the weightiest of all the conditions of progress in civilization and culture."[17] Before 1945 almost every scientist, even those working in Soviet Russia, took freedom in science, the right of an investigator not only to choose his problem but to attack it in any way that he could, for granted. Is this true today in Russia? It most certainly is not. Is it wholly true in this country? I hardly think so. This poses the question: How essential is complete freedom to a gifted investigator? There is no simple answer. There is no doubt of the necessity of freedom in science, but a discussion of how much involves questions that are bound up in the philosophy of science and the philosophy of the scientists. A discussion of this topic and others related to it by about 75 scientists from many different countries in a conference at Princeton University in 1947, brought out some interesting aspects of science, seldom mentioned in older treatises: (1) "The prosecution of applied science in its most efficient form is identical with that of pure science"; (2) "There is no scientific method as such, the most vital feature of the scientist's procedure has been merely to do his utmost with his mind, *no holds barred*"; (3) "We have been too complacent in the past in assuming that our diversities of opinion are not of fundamental significance. Diversities are symptomatic of something and we do not know what. It is time we analyzed our disagreements and found their significance. Have we a right to our bland assumption that the human race is intellectually all alike, or may there be truly irreconcilable points of view?"[18]

I should like to comment on the last of these statements, as it is so closely related to an understanding of the history of science. Modern science, as we know it and practice it today, had its origins in the late sixteenth century. But its roots go back to a much earlier period, to the intellectual achievements of the Greeks. Its development in the seventeenth century is a result of emphasis on reasoning in the late Middle Ages, and on the economic and social activities of the Renaissance. Our civilization is the result of forces which in most respects operated only in the countries of western and northern Europe, exclusive of Russia. Today whether we wish to do so or not, we must try to understand the motives and acts of Russia and the Orient in the light of the facts of their development and ours. Their opinions are not like ours. Is this difference of opinion of fundamental significance? Can peoples not conversant with our ideas, that even the man in the street takes for granted, namely, the objectivity of scientific truth and its foundation on abstract mathematical thinking, understand our way of life, our goals? Assuming that these peoples can eventually achieve an understanding of our civilization, will it come in one generation, or two, or many? The populations in the countries not within the sphere of western civilization far outnumber those of western Europe and the Americas. If, as I believe, the large majority of these peoples cannot understand our civilization based on the development of ideas of freedom, scientific objective truth, and on the technology which is an outgrowth of these, can we by study understand them? We most certainly must make the effort. But something more is needed than just the willingness to make the effort. We must understand our own civilization before we can learn much of that of another people.

Every graduate student specializing in chemistry is first and foremost a human being, a citizen of this country, responsible to an extent determined by his talents and ability to take part in preserving western

[14] WILLSTÄTTER, R., "Aus Meinem Leben," Verlag Chemie, Basel, **1949**, p. 324.

[15] HAUSER, ERNST A., "The lack of natural philosophy in our education," J. CHEM. EDUC., **28**, 492 (1951).

[16] STANDEN, A., "On the horns of the sacred cow," J. CHEM. EDUC., **28**, 608 (1951).

[17] YOUMANS, E. L., *op. cit.*, p. 262.

[18] "Physical Science and Human Values," Princeton University Press, Princeton, **1947**, pp. 140, 144, 151.

civilization. If he is an intelligent student he must inform himself in one way or another of the factors that have determined the development of this country and those like it, so that in the years to come he may exert his influence effectively to preserve what we think is valuable to us. It is my conviction that he can do this only by exercising his mind in thinking with the investigators and philosophers of the past, which means that he must know something of the history and philosophy of science.

This paper ends with this last sentence, but I cannot resist commenting on an article that appeared in the January number of THIS JOURNAL. It was written by several members of the Russian Academy of Science and sets forth the law which is to regulate the teaching and the acceptance of chemical structural theory as it has been conceived by the Central Committee of the All-Union Communist Party. "The decisions of the Central Committee to ideological problems and the sessions of the All-Union Lenin Academy of Agricultural Sciences have mobilized Soviet scientists for the solution of the problem of a critical analysis of the present state of theoretical concepts in all fields of knowledge and the struggle against the alien reactionary ideas of bourgeois science." The main part of the article is a diatribe against the theory of resonance and a glorification of the achievements of Russian chemists in the last eighty odd years. Much has been written lately against the soundness of the theory of resonance as it is used in the interpretation of the structure and reactions of organic compounds. Like other theories held in high favor years ago, certain features of it are likely to be retained. This is to be expected for the very essence of a theory is that certain aspects of it undergo modification. But the distortion of the facts of the history of theoretical organic chemistry in this article is amazing. Everyone will agree that Butlerov, Markovnikov, and Mendeleev were great scientists possessed of unusual scientific intuition, but this may also be said of Berzelius, Laurent, Couper, Kekulé, Kolbe, van't Hoff, LeBel, and others. The development of structural theory might have been delayed had these particular Russians, trained according to the standards in use in the schools of France and Germany, not entered the field, but to name them as the founders of structural theory is a rank distortion of the facts recorded in the literature of chemistry in the middle and late nineteenth century. One can be certain that if the history and philosophy of science are taught in Russian universities, western scientists would be astounded at the so-called facts recorded in the lectures and textbooks. The history and philosophy of science are international in character. Some scientists believe that their study is the last hope of peace among nations.

CHEMISTRY PROGRAM OF A. A. A. S. MEETING

THE program of Section C (Chemistry) of the A. A. A. S. meeting in St. Louis during the last week of December will consist of General Papers on Friday, December 26, the Celebration of the Centennial of Washington University on Monday, December 29, and excellent Symposia on Saturday, Sunday, Tuesday, and Wednesday (December 27, 28, 30, and 31).

The Symposia include titles such as: Contributions of Chemistry to New Textile Fibers, Contribution of Chemistry to Animal and Human Nutrition, Contributions of Chemistry to Medicine, Contributions of Chemistry to Dentistry, Chemical Economics of Fuels of the Future, Contributions of Chemistry to Agriculture, and an Alpha Chi Sigma-sponsored program on Professional Training.

Chairmen of some of the symposia include Dr. Frank Soday, Dr. B. L. Scallet, Dr. Barbara Campbell, Dr. Kenneth N. Campbell, Dr. Frances Krasnow, Dr. L. L. Hirst, and Dr. Virgil L. Freed. The symposia have been arranged by Dr. Carrol A. Hochwalt and Mr. J. J. Healey, Jr. both of Monsanto Chemical Company.

Titles and abstracts of papers for the general session on Friday, December 26, should be sent to E. F. Degering, Secretary, Section C, A. A. A. S., 1860 Colonial Drive, Memphis 7, Tennessee, *on or before the first of October.*

A banquet dinner for the chemists will be featured on Monday evening in connection with the Centennial Celebration of Washington University, at which Dr. Carroll A. Hochwalt, Chairman of Section C and Vice-President of A. A. A. S., as well as Vice-President of Monsanto Chemical Co., will address the group.

LANGUAGE TEACHING FOR STUDENTS IN SCIENCE AND TECHNOLOGY[1]

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EFFECTIVE salesmanship requires careful attention to the prospective customers' wishes, habits, and prejudices. Similarly maximum effectiveness in teaching is impossible without consideration of the student's background, motivation, and attitude. The salesman must often transmit information—"educate" his prospect—in order to motivate him to buy. The teacher, likewise, must be sure of the student's interest—must on occasion "sell" him—if effective teaching is to be accomplished.

This relationship between selling and teaching is particularly important when a foreign language is being taught to students of science and technology. Such students tend to be preoccupied with courses pertaining directly to their professional realm of specialization. Advances in science and technology during recent decades confront today's budding scientist or engineer with an overwhelming amount of material in his chosen field. The understandable and laudable desire to obtain the best possible training in his chosen field may result in something closely akin to sales resistance toward those subjects—such as foreign languages—which appear of little direct benefit to the student's future professional career. This sales resistance is encountered not only among students but, occasionally at least, among those charged with establishing curricula in the various fields of science and technology.

In meeting this situation, foreign language teachers often place too much reliance on one variation or another of the stock argument that work published in languages other than English cannot be neglected in a well-managed program of scientific research or industrial development. Even when the validity of this argument is conceded, students in science or technology may be inclined to the opinion that there is so much in English to be mastered that worrying about foreign language material is somehow a little bit ridiculous. And it is a fact that, in certain fields of science and technology at least, English is the principal language of publication (*1*). The present position of responsibility that the U. S. A. has been forced to assume in world affairs may even come to be misinterpreted as a reason for regarding the study of foreign languages as a poor investment of time and energy for future scientists and engineers. Nevertheless, in spite of misconceptions and prejudices, knowledge of one or more foreign languages can provide unique advantages in developing a career in science or technology. Discerning these advantages is often difficult for the student at the start of his career.

A very important—perhaps the most important—factor in this situation is the attitude of the foreign language teacher. He must be willing to adjust the objectives of his courses and his teaching methods to the requirements of the students. Their interest in language study will rarely go so far as a desire to acquire the ability to speak a foreign tongue. A much more realistic goal—at least initially—is the ability to read foreign language material in the student's chosen field—be it physics, electrical engineering, chemistry, mathematics, or the like.

Failure to establish reasonable goals in teaching foreign languages to students in science and technology has been a principal cause for the existing chaos in standards and also in teaching methods (*2*).

The definition of a reasonable initial goal is, of course, only the first step. Teaching methods appropriate to the goal require careful attention. The teacher must understand what the student must learn in order to achieve an effective reading knowledge which will be useful to him in his subsequent professional career. Perhaps the most exasperating mistake which the teacher can make in this connection is to choose the wrong type of material for practice reading. Students in science and technology are quick to see that the vocabulary required for reading fairy stories is quite different from that needed to read papers dealing with electronic circuitry, chemical reactions, or suspension bridges. Reading material in the student's own field of specialization can be used right from the start, even in beginners' courses (*3*). For example, with chemists tackling Russian for the first time, it is well to use an elementary inorganic text such as Verkhovskii's "Neorganicheskaya Khimiya" (*4*).

Important advantages are gained by using foreign text material whose subject matter is well known to the student. With such reading material, a knowledge of basic chemistry—or other field of science or technology—can assist the student very effectively in acquiring insight into the mechanism of expressing thought in the foreign language. Another advantage is the feeling of accomplishment that the student experiences early in the course. In order to maintain this feeling, once a minimum level of proficiency has been attained, it is well to shift to reading material whose subject matter, though in the student's field, is not entirely familiar to him. By using a little care in selecting reading mate-

[1] Presented at the XIIth International Congress of Pure and Applied Chemistry, New York, September, 1951.

[2] Present affiliation: Bjorksten Research Laboratories, Madison, Wisconsin.

rial of gradually increasing difficulty, the student can be led by stages to reading papers reporting current advances in his field (*5*).

In addition to skillful selection of appropriate reading material the teacher needs to devote careful attention to his presentation of grammar fundamentals. There is danger of giving the impression that the language being studied is an impossibly difficult puzzle, characterized by a multiplicity of confusing rules and irregularities. Perhaps the most effective means for avoiding this pitfall is to emphasize the general structure of the language being studied and approach the finer points of grammar by considering example sentences. The student should be given the impression that grammar rules are analogous to a road map which is very helpful in becoming acquainted with new territory, but which has served its usefulness once familiarity with the strange landscape and its land marks has been achieved. Sentence structure should be emphasized rather than grammar rules. Example sentences from scientific or technical material are particularly helpful in this connection (*6*).

Another point that the teacher should keep in mind is the fact that foreign language courses may require the student in science or technology to develop study methods different from those he ordinarily uses in coping with subjects whose mathematical basis requires primarily skill in discerning which fundamental principles, and corresponding mathematical formulas, are to be applied to a given problem. In contrast, building up a network of associations as an aid to the memory is the basic technique in language study. This is particularly true in acquiring effective vocabularies. Here many students blunder—without realizing it—into a dangerous pitfall. Since dictionaries present arrays of foreign words with one or more English words placed opposite a given individual foreign word, there is a tendency to assume that one, or possibly two, English words are fully and precisely equivalent to the given foreign word. There is a tendency, in short, to set up equations and to attempt to understand the foreign language by a mechanical process of replacing the foreign words with what are presumed to be English equivalents. It is true, of course, that with words designating certain simple material objects, something close to a true equivalence of meanings of a foreign and an English word may well occur. Such cases of equivalence must not be regarded, however, as typical.

The student must be warned in good time that the meaning of a foreign word must be expected in the majority of cases to differ more or less extensively from the meaning of any one English word. The student must be taught to use the dictionary as a means for getting acquainted with the foreign word. He must come to realize that thorough understanding of foreign words and foreign phraseology can be achieved only by careful observation as to how words are actually used.

This process of becoming aware of the exact area of meaning of a foreign word is best acquired by practicing the habit of reading without translation. Unfortunately, in classroom recitation, translation is the only method for testing and confirming that the student—particularly the beginner—is understanding a passage in a foreign language. In spite of this necessity of resorting to English in the classroom, the student—even the beginner—should be urged to acquire the habit of associating a semantic impression directly with the foreign word without interposing English in the understanding process. In order to promote this habit of understanding a foreign language directly without translation, it is particularly important to make clear to the students the futility—even harmfulness—of resorting to writing English words in a foreign text so as to provide an interlinear translation. If difficulty is experienced in remembering the meaning of a foreign word, it is much better to use separate slips of paper which can be discarded in favor of direct understanding as soon as the foreign word has been firmly anchored in the memory.

Over and above the organization of the course and assistance in developing effective study methods, the teacher can do a great deal in direct "selling" of foreign language study to the student. Perhaps the most effective sales approach with students in science and technology is to bring them to the realization that scientific and technical literature in their fields is not entirely in English and that knowledge of a foreign language may serve as a key to much valuable knowledge. It is perhaps a little more difficult to convince students of science and technology that there are certain accessory benefits to be gained by foreign language study. Nevertheless, these benefits do exist and it may not be out of place to mention some of them briefly.

Acquiring insight into a foreign language provides a surprisingly effective basis for understanding one's own language, and thus acquiring skill in the difficult art of expressing ideas clearly and convincingly. In these days, the ability to prepare well-written reports and papers that are effective in selling ideas may prove to be the decisive margin of success in developing a professional career in science or technology. It is true, of course, that such skill can be developed without learning foreign languages, but this does not detract from the fact that successful mastery of a foreign language is a very effective means for developing the ability to write effectively.

Another secondary benefit in learning foreign languages is the possibility of easier and more pleasant communication with foreign colleagues. It may be thought that it is necessary to acquire a complete speaking mastery of the foreign language in order that communication with foreign colleagues may be facilitated. Students need reassurance on this point. They need to be brought to understand that in dealing with foreign colleagues face-to-face, a knowledge of the foreign language, far short of mastery, may prove extremely helpful in achieving communication and understanding. By knowing something of the structure of the other man's language, his attempts at speaking English become much more easily understandable. By knowing something of how his language sounds, his pronuncia-

tion of English will become much less strange. Finally, if, on occasion, it is possible to help the foreign colleague by a few phrases from his own language, it is easy to give him the impression that an attempt is being made to meet him part way. This is helpful in establishing good personal relationships on a friendly basis.

Many young engineers and scientists may find it hard to believe at first that getting a view of another pattern of expressing thought can be an interesting and satisfying experience. If the teacher of a foreign language is really skilled in his job it is possible to present a foreign language in such a fashion that the student of science or technology may become fascinated with it as a mechanism for communicating thought. Scientists and technologists are by vocational aptitude interested in devices in the physical and material realm. It is possible to arouse their interest in those intellectual tools which languages have developed to express ideas and emotions. It is even possible on occasion to bring the young scientist or engineer, whose interest in language initially was either utilitarian or perhaps compelled by curriculum requirements, to an understanding that the literary masterpieces of a foreign tongue can provide thrilling experiences. Observing this development in one's students is, however, frequently denied the teacher as such realization generally comes long after graduation.

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THE BOSTON COLLEGE CHEMICAL SOCIETY

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THIS year we celebrated the 25th anniversary of the organization now known as the Boston College Chemical Society. For the occasion, the student members did some historical research and published a brief history of the activities of this campus group. It tells the story of the ups and downs of an organization that is subject to a rapid turnover of membership and officers, now surging forward on waves of enthusiasm, now falling back into the trough of the wave. The history relates that at one period it submerged and operated in submarine quarters with such officers as the Grand Molecule, the Grand Atom, and the Grand Ion. Eventually it resurfaced, and after the war years, although it is still subject to the wave phenomenon, we have been riding the crest of the wave with more success since our affiliation with the American Chemical Society.

Our organization must be such that it makes membership worth while to the busy chemistry student, and affords opportunity for active participation by all. The function and activities of our group are largely determined by the circumstance that Boston College is predominantly a college of commuters. About five years ago we undertook a common project of studying parliamentary procedure with a view to formulating and adopting a set of bylaws. This project was not only a profitable study for future professional men but it also furnished us a student-made guide, tailored to our campus picture, and sufficiently broad in its objectives to allow a variety of activities while keeping within reasonable bounds. Article II of the bylaws reads: "The object of this Society shall be to afford an opportunity for the students of Boston College interested in chemistry to become better acquainted, to secure intellectual stimulation that arises from professional association, to secure experience in presenting technical material before chemical audiences, to foster a professional spirit among the members, and to instill professional pride in chemistry."

From year to year there is considerable variation in the group spirit and group leadership with corresponding variation in activities and emphasis on the above objectives, but three annual events have become traditional by success, and they are planned well in advance.

(1) Early in the fall, while the weather is still mild, an outing is held at a readily accessible State Reservation where good recreational facilities are available. The highlight of the sports events is a game of softball between the undergraduates and the graduate students. A "hot dog" roast with songs around the campfire, and into the darkness, concludes this get-acquainted activity. (2) Toward the close of the football season a smoker is held, at which a popular feature is the showing of football films with commentary by a member of the coaching staff. Mimeographed song sheets help to promote community singing to complete a pleasant evening. (3) In the spring a somewhat elaborate banquet is held at a downtown hotel, and despite the heavy expense our banquets have been well attended. Each year these events are voted by the membership, all arrangements are made by elected committees, and it would seem that they effectively serve the purpose of unifying the various elements of the group.

For some five years before the war we published a quarterly bulletin called "The Crystal." Enthusiasm ran high and the quality was good for two or three years, but succeeding classes did not feel the enthusiasm of the founders, the quality dropped gradually, it became more difficult to meet the publication deadline with an issue that was worthy of the expensive format, and in its weakened condition it became a ready war casualty. Instead of reviving "The Crystal," as some enthusiasts proposed, we started a less pretentious monthly organ called the "Chem Bulletin," edited by the students and lithoprinted at the College Press. It is intended for local readers and it is well received by them. A typical issue contains news of the chemistry department, introduction to a faculty member, editorial on the morality of using the H-bomb, a discussion of the qualities required in a chemist, notice of graduate seminars, an article on job opportunities, news of the alumni, a crossword puzzle, and a humor column. This activity is now in its fourth year and it is still vigorous. It is a sizable job for the editor, but he has a large editorial staff, and at his discretion he may omit an issue if it would seriously interfere with studies. Technical articles on an undergraduate level are welcome and they find a place in the "Chem Bulletin" with sufficient frequency to justify the title. A sponsor is defined as one who makes himself responsible for another, and, naturally, the Faculty Sponsor of the group is responsible for the irrevocable printed work. In reviewing the copy of our publication his work is restricted largely to the detection of errors of judgment, and with an established editorial policy that nothing may be printed that could cause offense to any individual or institution, his work is relatively easy.

During the school year we manage to have about 12 regular meetings, of which four are scheduled for business, two for science films, and six for speakers. The supply of willing and competent speakers in our area is more than adequate, yet the program committee finds it difficult to obtain speakers who have something of general interest and within the easy understanding of a group composed of sophomores, juniors, and seniors. This year's sophomores will be next year's juniors, and if they feel out of place and neglected their loss of enthusiasm will be reflected in the organization for the next two years. Hence, we try to avoid highly specialized topics that would be suitable for a graduate seminar or for a selected group of seniors. Job opportunities; description of a local chemical industry, personal experience of a fairly recent graduate, the chemist in civil service, the police chemist, etc., seem to be attractive topics for the purpose as measured by attendance. Strangely enough, our most successful speakers have been student members who had either gained some experience in summer work or who had developed an interest in some special field of chemistry by study or by working at a laboratory project. Unfortunately, we have had too few volunteers during the past two years to constitute a significant part of our program.

This leads to the mention of another activity that is now in the trough of the wave. Group laboratory projects according to interests were proposed, modest space and equipment were provided, and cooperation of the faculty was assured for direction of the projects. Real enthusiasm was shown for two years in projects such as the isolation of natural products, clinical tests, glass blowing, etc., and some of the participants gave excellent talks at the regular meetings. At present a few students are working on individual projects, but group work has disappeared—causes unknown. However, the opportunity is still open and quite possibly the activity will be revived with a new group next year. The annual variation in interests and leadership of the group is amazing.

Another sporadic activity is the visitation of local industrial plants and laboratories. Such plant trips must be made during school vacations, and are arranged by small groups, who must guarantee to the faculty sponsor that they will arrive as a group at the plant or laboratory at the time specified by the host. Usually two or three such plant trips are made annually.

A student organization that has recently assumed a role of some importance in this area is the Intercollegiate Chemical Society, which draws its membership from about a dozen College Chemistry Clubs in the neighborhood of Boston. Our group has lent its support to and has actively participated in the program of this relatively new and active organization.

As an illustration of how an active chemistry club can bind the members together in ties of friendship and professional interest, with mutual benefit to the members and the alumni, several of our members undertook to assemble and publish an alumni directory. This project arose from a desire to know what their friends of last year were doing, and it matured into a complete directory that was mailed to the alumni last year. This will be a continuing project of the group.

Affiliation with the American Chemical Society undoubtedly has had a beneficial effect on the group, but it is difficult to measure even qualitatively the

direct influence of this affiliation on any one activity. Rather, it is an over-all effect in producing those real but intangible qualities called professional pride, maturity of outlook, and professional attitude. The feeling of belonging to a large professional organization lifts them out of their campus isolation, and also imparts a sense of solidity to their local activities. Only on this basis can one explain a spontaneous movement two years ago that resulted in a migration of 25 members to Philadelphia to attend the national convention of the A. C. S. They wanted the experience of seeing their organization in full operation.

In closing, the observation may be made that by active participation in the activities of his Chemistry Club a student gives something and receives something. The rule still holds that he receives as much as he gives.

PRACTICAL PRECISION FOR CHEMISTRY STUDENTS[1]

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PRECISION of measurement has occupied my interest and attention at irregular intervals for the past 30 years. During this time I have been inclined to attach more and more importance to the subject and to teach more and more of it to students of chemistry. Perhaps as a teacher I may be likened to a father who is bent on giving to his sons the advantages and opportunities that he did not have when he was a boy. I trust that I may be pardoned, then, if I appear to overemphasize the importance of precision of measurement and the need for more training in this subject in our collges and universities. My early struggles with precision are still vivid in my mind. I graduated from college with just enough misinformation to be impelled to make all measurements from 10 to 20 times as precisely as necessary. I can recall many hours spent regulating the temperature of a water bath to the nearest thousandth degree, hours which were wasted, as I long after realized, since regulation to the nearest hundredth degree easily satisfied the precision requirements of the measurements.

Out of a spirit of humility I might be inclined to believe that my early blunderings in precision of measurement were unique; however, the experiences of three decades have convinced me that a lack of a practical knowledge of this subject is quite common, I might even say is almost universal, not only among young graduates but also among those skilled and experienced in carrying out experimental work. One continually encounters cases of neglect of precision in the work of graduate students, in the accounts of research in reputable scientific journals, and even in the design of apparatus furnished by makers of scientific equipment. Let me mention only two or three examples to illustrate.

In one report on an investigation in which slight changes in concentration of an acid were to be determined rather precisely, an ordinary 1-ml. volume pipet was indicated as a means of taking samples for titration against a standard base. As a result, the uncertainty in the determination of concentration of the acid was hardly less than the change in concentration to be determined. Conclusions reached on the basis of these measurements were perfectly worthless.

In a master degree thesis a student reported the determination of the ammonia content of a sample by distilling it into a large excess of 0.1 *N* acid and back-titrating with 2 *N* standard base. A rough precision analysis quickly exposes the fact that the practical uncertainty in the result is of the order of magnitude of five per cent although the student reported the result with an apparent precision of one part in a thousand.

For a considerable time there have been manufactured and sold so-called "microburets" which are graduated in 0.01-ml. divisions and which may be read to a decimal precision of 0.002 ml. Many of these burets are designed so that they furnish from their tips drops that are 0.02 ml. or more in volume. Apparently it has been overlooked that a titration normally is only as decimally precise as half the volume of the drop falling from the tip, no matter how precisely the reading of the buret may be taken. Fortunately this anomaly in design has been corrected recently

[1] Presented before the Chemistry-Physics-Mathematics Section at the annual meeting of the Northwest Scientific Association, December 28, 1951.

and there are now available buret tips that furnish "microdrops" of a size corresponding to the precision with which the buret may be read.

After one has decided that training in precision is an important requirement for every student of chemistry, several questions immediately present themselves. When should this training be given? Should it be given in the department of mathematics or in the department of chemistry? Should it comprise a separate course or should it form part of a required course such as analytical chemistry or physical chemistry? The answers to these questions will probably depend to some extent upon the particular curriculum at any given institution, but generally, I believe, they will follow along much the same lines.

Precision of measurement is not what students would call a "pipe course." It continually involves keen analytical reasoning as well as a knowledge of measurement and an understanding of differential calculus is a prerequisite. The attempts that I have made to give small doses to sophomore students taking analytical chemistry have not been particularly successful. The subject is definitely an "upperclass" one and should not be given, I should say, until the junior year.

Precision of measurement is an extremely practical subject and is involved every time any reading—temperature, pressure, buret, colorimeter, and the like—is taken. It would appear that one can emphasize the practicality of the subject as well as illustrate its applications much more effectively when the subject is given as a part of a course in which measurements are made, and in which measured quantities are used in mathematical operations and involved in various mathematical relationships. In the chemistry curriculum the course in physical chemistry naturally comes to one's mind in this connection. The principles may be taken up as part of the lecture program with examples taken from the experimental work of the course. Practice in the application of the principles may then be obtained by keeping the student "precision minded" at all times while he is carrying out his laboratory assignments.

I presume that one can safely predict that many teachers of physical chemistry will strongly object to crowding any more material into an already crowded course in physical chemistry. They will argue that one cannot do justice to the field of physical chemistry in the regular course of five hours for one year, and that material on precision can be included only by crowding out something else. That is all quite true. Even without the inclusion of precision measurement one cannot cover thoroughly all of the material included in the average text of physical chemistry. On the other hand, much of the material ordinarily considered in the first course in physical chemistry may be classed as "applied" and is often so superficially treated as to confuse the student with a mass of poorly learned facts and half-understood relationships. The question is one of relative importance. Discussion of subjects such as quantum mechanics and nuclear reactions and the like, I believe, may well be postponed to a later course and the time so saved spent to better advantage in teaching the principles of precision.

For some years at Seattle University we have been including in the general course in physical chemistry a short but rigorous treatment of what one may call "practical precision." The objectives are (1) to train the student in methods of precision analysis, and (2) to develop in the student the habit of making a preliminary precision analysis every time he is required to make any measurement. Class discussion for the first three weeks (nine class periods) is devoted exclusively to this subject. The formal class work is then discontinued but the student is not allowed to drop the subject abruptly at this time. In the laboratory throughout the year the importance of precision is emphasized by the instructor and the student is expected to have an understanding of the precision relationships involved in any experiment at hand. Also, a complete precision analysis is a required part of many of the final laboratory reports. In this way the student has the opportunity of obtaining experience in and developing the habit of precision analysis to such an extent that we trust, every time he makes a measurement of any kind, he asks himself: (1) "How closely can I take this reading?" and (2) "How closely should I take it so that the requirements of practical precision are met?"

TABLE 1
Outline of Course in Practical Precision

(1) *General Principles and Definitions*
- (*a*) Errors and uncertainty in measured quantities
- (*b*) Significant figures
- (*c*) Certain and uncertain figures
- (*d*) Precision and accuracy

(2) *Precision Relationships in Mathematical Operations Involving Measured Quantities*
- (*a*) Multiplication and division
- (*b*) Addition and subtraction
- (*c*) Mixed operations
- (*d*) Operations involving logarithms
- (*e*) Operations involving trigonometric functions

(3) *Practical Precision Analysis in Experimental Work*
- (*a*) General methods of precision analysis
- (*b*) Outlines of precision analysis for selected experiments

The outline of the course is given in Table 1. The first section follows pretty much the traditional method of approach, but the second and third sections depart considerably from the traditional treatment. Consequently, the choice of textbook presented a problem in the beginning. Of the several texts available, not one was found to be satisfactory for the purposes of the course. Over the years there has been worked out a set of mimeographed notes, which at present has attained a size of 40 single-spaced typewritten pages. This appears to satisfy the requirements of the course as far as text material is concerned.

The first section opens with a discussion of the various kinds of determinable errors and undeterminable uncertainties associated with the taking and recording

of measured quantities. This is followed by the definition and elucidation of "significant," "certain," and "uncertain" figures, after which the terms "relative precision" and "decimal precision" are defined and a distinction is made between precision and accuracy.

The second section is concerned with relationships between the precision of one or more measured quantities and the precision of the result obtained by a mathematical operation involving the measured quantities. Considerable attention is given to operations involving multiplication and division. The "additive rule" is demonstrated using algebraic and geometric methods as well as the methods of the calculus. This is followed by the logical reduction of the rule to the simplified form: *In operations involving only multiplication and division the relative precision of the result may be considered to be identical with that of the least precise factor involved.* Upon the basis of this simplified rule the relative and decimal uncertainties associated with the result may be calculated and the result written with the correct number of significant figures. Students encounter little difficulty in understanding and applying the above rule to the solution of problems. However, the process involves two separate mathematical steps and is altogether too cumbersome for many of the requirements of precision analysis. It is necessary to devise a simplified method for determining the number of significant figures in a result, one that involves only mental arithmetic on the part of the student. The rule most often seen in books of quantitative analysis is that one should retain the same number of significant figures in the result as is found in the least precise of the measured factors involved. It is easy to demonstrate that this rule is very inaccurate in many cases and it is difficult to understand how it has been so widely adopted. Just one example should be sufficient to discredit it. According to the rule the result of the quotient 99.8 ml./97.4 ml. would be written 1.02. This result would have associated with it an uncertainty of 0.5 per cent, which is approximately ten times the uncertainty associated with the original factors.

What is needed is a method both precise and simple and one that can be easily rationalized for the students. The following rule covers both the determination of the number of significant figures and the estimation of the uncertainty in the result. It may be stated in two parts as follows:

(1) Continue to write figures in the result until there is obtained a number the size of which is at least as large, *decimal points excluded*, as the least precise of the measured quantities involved.

(2) With a given relative uncertainty, the uncertainty in the final figure of a number is directly proportional to the size of the number, *decimal points excluded*. The uncertainty in the final figure of the result may be obtained, therefore, by a direct proportionality involving the uncertainty of the least precise measured quantity, the measured quantity itself, and the final result, *decimal points excluded*.

To illustrate the use of the rule we may consider the problem of calculating the area of a rectangle the sides of which have been measured as 39.6 cm. and 0.16 cm. Performing the multiplication without consideration of the precision relationships we may write,

$$(39.6 \text{ cm.})\,(0.16 \text{ cm.}) = 6.336 \text{ sq. cm.}$$

The factor 0.16 is relatively the least precise of the two and so determines the precision of the result. The first figure of the result is 6 which is not as large as the number 16. Including the second figure we obtain the number 63, which is larger than 16. The result, therefore, is written with two figures. The number 63 is four times the number 16; therefore, the uncertainty in the final figure 3 of the result is four times that of the final figure 6 in the original factor.

$$\text{Uncertainty in } 3 = \frac{(63)}{16}\,(\text{uncertainty in } 6)$$

$$\text{Uncertainty in } 3 = (4)\,(0.5) = 2$$

Therefore:

$$(39.6)\,(0.16) = 6.3 \pm 0.2 \text{ sq. cm.}$$

At first sight the process may seem to be about as complicated as that described previously but actually it is not. After a little experience the student usually is able to carry out the calculation mentally almost as rapidly as the result may be written down.

With operations involving multiplication and division covered the student is introduced into more complicated operations, including those involving logarithmic and trigonometric functions. This material is absorbed without difficulty by those students who have a good knowledge of differential calculus.

The third section of the course includes the practical application of the principles of precision to actual problems of measurement. The general methods may be illustrated by the analysis of the determination of the molecular weight of a liquid by the well-known method of Victor Meyer. In this determination a sample of a liquid of weight g is vaporized in a bulb kept at constant temperature and the volume V of air displaced is measured at room temperature T and pressure P. The molecular weight of the liquid is calculated by means of the expression,

$$M = \frac{gRT}{PV}$$

where R is the gas constant, the value of which is known with sufficient precision. The several steps in the pre-

TABLE 2

Precision Analysis of Molecular Weight Determination by Victor Meyer Method

Factor	*Approx. size*	*Best decimal precision*	*Relative effect on result, %*	*Practical decimal precision*	*Practical least count*
g	0.2 g.	0.0001 g.	0.05	0.0006 g.	0.001 g.
P	740 mm.	0.05 mm.	0.007	2 mm.	1 mm.
T	300°	1°	0.3	1°	1°
V	40 ml.	0.05 ml.	0.13	0.12 ml.	0.1 ml.
Over-all Practical Precision 0.3%					

cision analysis are indicated in Table 2. In column one are listed the four measured quantities, the approximate sizes of which are listed in column two. In the third column are shown the best decimal precisions possible with the instruments available and under existing experimental conditions. It will be noted that the best decimal precision of the reading of temperature is not determined by the precision with which a thermometer may be read but by the precision with which the temperature of the air surrounding the apparatus may be regulated. In the fourth column are listed the relative uncertainties produced in the calculated value of the molecular weight as results of the uncertainties in the several measured quantities. The largest of these, 0.3 per cent, fixes the over-all practical precision of the determination.

In column five are shown the practical decimal precisions with which the several factors are to be measured so that they may correspond to the over-all practical precision, 0.3 per cent, while in column six appears the corresponding least counts with which the measurements are to be made. In selecting the practical least count a precision slightly better than the over-all practical precision is used whenever possible to prevent the accumulation of unnecessary uncertainties in the final result.

The precision analysis of experiments involving measured quantities and final results in simple linear relationships is not too difficult for the average student of physical chemistry. Unfortunately, many of the determinations involve mixed relationships including addition and subtraction as well as exponential and trigonometric functions. Experience has demonstrated that the average student is unable to carry out a precision analysis of such determinations without assistance. Apparently he lacks sufficient experience to know how to begin, what assumptions can be made, and for what to solve. The last feature of the course, a series of outlines of precision analysis for the common experiments of physical chemistry, is designed to assist the student in preparing precision analysis reports for determinations as he encounters them in the laboratory. These outlines include a minimum of necessary information as well as hints and suggestions to enable the student to work out the precision analyses. The outline for the precision analysis of the determination of concentration with the photoelectric colorimeter shown below is representative of the outlines in this series. During the course of the year the student is required to work out and hand in at least ten of these precision analyses as parts of the final laboratory reports.

The results obtained from the work in precision have been generally quite satisfying. The students soon become "precision minded," so that the main objective of the course is realized; and many are prepared to undertake almost any problem in precision analysis that may be presented. In addition, there appears to be an unexpected dividend in the form of a pronounced increase of interest in the experimental work. The students are more interested even in those experiments that are usually classed as dull or boring, an interest that originates apparently from a desire to obtain results with a precision that approaches that which has been calculated. The combined results, I believe, justify the introduction of the material on precision into the course in physical chemistry.

Determination of Concentration with the Photoelectric Colorimeter

The relationship involved in the measurement of concentration with the photoelectric colorimeter is the Lambert-Beer law.

$$\text{Log } I/I_0 = -kc, \text{ or } c = (-1/k) \log I/I_0 \quad (1)$$

The colorimeter has two scales, the T scale and the A scale. The per cent transmission "T" is defined by the relationship:

$$T = 100\, I/I_0$$

Substituting this in equation (1) we obtain the relationship between the concentration and the reading on the T scale:

$$c = \frac{2 - \log T}{k} \quad (2)$$

The unit of the A scale is defined by the relationship:

$$A = -100 \log I/I_0$$

Substituting this in equation (1) we obtain the relationship between the concentration and the reading on the A scale:

$$c = \frac{A}{100k} \quad (3)$$

Precision Analysis. The function of a precision analysis for this experiment is the estimation of the relative precision obtainable in the determination of concentration by means of the photoelectric colorimeter. To accomplish this it is necessary to develop the relationship between the precision of the scale reading and that of the concentration.

(*a*) Note which scale on the colorimeter has even divisions. The decimal precision of the reading is the same at all points on this scale.

(*b*) Select the proper equation from those above and solve for dc/c.

(*c*) Note the decimal precision with which the scale may be read and calculate the relative precision with which the concentration may be determined at several points on the scale.

(*d*) Determine the point of finest precision using the method of maxima and minima.

(*e*) Construct a curve and locate the boundaries of the region within which a precision of 0.5 per cent may be obtained.

JAN INGENHOUSZ

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MENTIONED among the "forgotten chemists" by the late historian of chemistry E. F. Smith,[1] Ingenhousz, in more recent times, has been recognized by photo- and biochemists as an eminent early worker in the field of photosynthesis. His most important contribution to science "Experiments upon Vegetables," Emsley, London (1779) has just been re-edited and annotated by H. S. Reed.[2] His life story has been recorded by three of his compatriots, a physician, a botanist, and a physicist in 1875, 1880, and 1897, and lastly by an Austrian botanist in 1905. A contemporary of Priestley, Scheele, and Watt, he was 20 years older than Lavoisier but considerably younger than Sir John Pringle (1707–82) and Franklin (1706–90) whose friendship he greatly valued.

Ingenhousz was born on December 8, 1730, in the small town of Breda, known in history as the place where the third war between England and the Dutch republic was terminated in 1675 with the loss of New Amsterdam to the British. He got his early training from his father Arnold Ingenhousz, a physician, his mother having died the year after his birth. Having received a classical education at the Latin school of his home town he enrolled, at the age of 16, at the nearby Catholic University of Louvain, to study medicine. Here he obtained his M.D. at the age of 22, whereupon he went to Leiden for additional training in physics under the renowned Petrus van Musschenbroek (1692–1761) and in chemistry, as taught by David Gaub (1704–1780), pupil and successor of Boerhaave. After brief excursions to Paris and to Edinburgh for special courses in obstetrics, Ingenhousz settled as a practicing physician in Breda for eight years (1757–65) and kept up his interest in natural science by home study and correspondence with leading scientists in the Netherlands and elsewhere.

Realizing that for a Catholic there was no future in academic milieus in his native land, Ingenhousz left Breda shortly after his father's death to establish a medical practice in London. He took this step at the urgent solicitation of Pringle, who as a military surgeon with the English army from 1742–48 had repeatedly passed through Breda and had become acquainted with the gifted young Dutchman. In London Ingenhousz took an active part in smallpox inoculation as it was carried out in England. On Pringle's recommendation Ingenhousz was appointed court physician in Vienna. On the way over to Austria he called on the Austrian ambassador in the Hague who reported to his superiors that "Mr. Ingenhousz seems to be very intelligent, well versed in the sciences and yet a very modest man."

The coming of Ingenhousz to Vienna was due to the fact that the Empress Maria Theresa had lost several relatives who were victims of smallpox and hence was very anxious to protect her still untouched children from this dreaded disease. In Vienna, where he was given the princely salary of 5000 florins, Ingenhousz was received with open arms by his fellow Dutchman Gerhard van Swieten (1700–72), likewise a court physician who had left his native land in 1745 because as a Catholic he was unacceptable to succeed his teacher Boerhaave.

Ingenhousz was successful in the inoculation of the Empress' children and the reputation thus acquired played an important role in securing him the leisure required to pursue his avocation as a student of nature. In the introduction to his first major publication he writes:

> The study of nature does not enrich the observer. If I were looking for gain, I would have abandoned this pursuit to follow the path of fortune which fate had opened to me. The work of the spirit cannot be forced on us like manual labor.

Ingenhousz' life as a scientist would have been rather dull if it had not been for his insatiable desire to travel which brought him much-needed distraction from his labors as a naturalist. He was very fond of England which, at that time, was very productive in scientific achievements. In 1769 he visited Tuscany, and the next year Elba, Switzerland, Paris, and Holland. In 1771 he visited his old friend Pringle who introduced him to Franklin, at that time agent in London for the colonial assemblies of Pennsylvania, Georgia, New Jersey, and Massachusetts. Together with Franklin and his grandnephew and an English friend John Canton, Ingenhousz traveled to the north of England.[3]

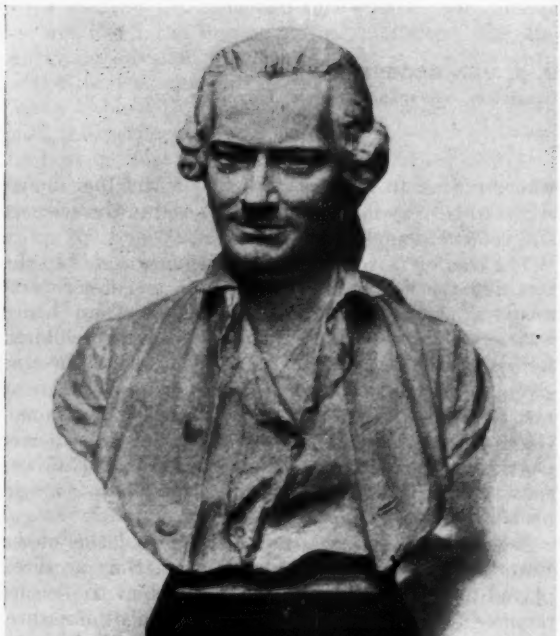
In 1772 Ingenhousz returned to Vienna. In 1775 he married the sister of his Dutch colleague von Jacquin, a professor of botany at the University of Vienna. The problem on which Ingenhousz had been working since 1773 finally found a solution in 1778 on another visit to England, where he had the time and the means to make an exhaustive study of the relations between plants and the atmosphere. The problem is well stated in Ingenhousz' own words:[4]

[1] SMITH, E. F., J. CHEM. EDUC., **3**, 29 (1926).

[2] REED, HOWARD S., "Jan Ingenhousz, plant physiologist," *Chronica Botan.*, **11**, 285–396 (1949).

[3] Carl Van Doren in his "Benjamin Franklin's Autobiographical Writings," New York, **1945**, p. 208, states that no record of this important journey has been left except in a brief letter of Franklin to his wife.

[4] REED, HOWARD S., *op. cit.*, p. 348.



Jan Ingenhousz

If the faculty which plants possess of yielding dephlogisticated air, of correcting foul air, and improving ordinary air, depended upon their vegetation, they would exert it at all times, and in all places in which vegetation goes on. A plant may vegetate and even thrive very well, in the utmost darkness; and yet in such a place it has no power to correct bad air, or to yield good; but on the contrary, it spreads round about it deleterious exhalations, which render the best air even pernicious to the utmost degree.

It will not be difficult to understand now from what cause all those different and contrary effects which Dr. Priestley has found in his experiments did really depend; and why Mr. Scheele had constantly found that the vegetation of beans always spoilt good air.

Priestley had previously performed some experiments on a cucumber placed in a bell jar full of air which had become vitiated by the respiration of animals, and had found that when the enclosed gases were tested they were again fit for respiration and supported combustion. However, he could not always repeat this experiment. Sometimes the air was purified, sometimes it was not. Ingenhousz found the solution by showing that only the green parts of plants purify the air, and only in sunlight. Like a great many early philosophers he assumed that the process was designed primarily for the respiration of animals, more particularly for the climax of creation, man himself. That plants, especially those in flower, in the shade secrete "concealed poison" was known to Ingenhousz although he did not realize that this poison (CO_2) arises from the process of respiration, common to all living cells, both plants and animals. The conclusions which Ingenhousz reached in his "Experiments upon Vegetables" involved him in heated controversies not only with Priestley, who evidently did not differentiate between the two processes of gas evolution, but also with another noted clergyman and amateur-scientist, Jean Senebier of Geneva (1742–1809). The latter carried out extensive experiments on the influence of sunlight, particularly on plants, which were published in four volumes. Plotnikow in his "Allgemeine Photochemie" (Leipzig, 1920) completely ignores Ingenhousz and calls Senebier one of the most important and talented representatives of pure photochemical research. On the other hand, Dhar in his "Chemical Action of Light" (London, 1931) gives due credit to Ingenhousz' discovery.

Ingenhousz' scientific achievements, which were rewarded in 1779 by his election as Fellow of the Royal Society, extended also to the rapidly advancing fields of physics and chemistry. In 1778 he discussed before the Royal Society Franklin's theory of positive and negative electricity, and the next year in another lecture described an improved apparatus for generating frictional electricity. Noting the close analogy between heat conductance and electrical conductance he set up an apparatus for measuring the heat conductance of various metals. For this work he tried to enlist the cooperation of Franklin, at that time (1780) residing in Paris. Lack of time prevented the execution of this plan and Ingenhousz returned to Vienna to carry out the work alone. This research, for which he gave Franklin full credit as originator, was later fully confirmed by Wiedemann and Franz.

Little needs to be said of the purely chemical experiments performed by Ingenhousz, such as the combustion of camphor, phosphorus, iron, and steel in what was then called "dephlogisticated air," in which he took great delight. The types of apparatus which he used were all of the conventional kind and showed no special features.

A number of his miscellaneous papers on physical and medical topics were published in Vienna in 1784 and the next year a volume on "Further Experiments in Physics" was issued in Paris with a dedication to Franklin. The relations between Ingenhousz and Franklin were very close at that time. In a letter of January 14, 1785,[5] Franklin writes:

. . . and let me know soon if you will make me happy the little remainder left me of life, by spending the time with me in America. I have instruments, if the enemy did not destroy them all and we will make plenty of experiments together.

The library of the American Philosophical Society in Philadelphia contains a few Franklin and Ingenhousz items which are illuminating. Franklin, in his capacity as ambassador to France, wrote, previous to his return to America in July, 1785, the following passport:

. . . to all commanders of armed vessels, to let Dr. Ingenhousz go on his private affairs, to the United States to return to the continent within six months, and as he is a person of distinguished

[5] A page from this letter is reproduced in *Chronica Botan.*, 11, 380 (1949).

merit, long known to me and my intimate and dear friend, I recommend him particularly to your civilities, assuring you that I shall esteem and acknowledge them as done to me.

Ingenhousz, through his relations with Franklin, had been in touch with a Mr. Wharton to whom he had delivered £8000 for investment in the United States. It seems that this sum of money was a total loss for in Franklin's last letter to Ingenhousz dated February 11, 1788, we read:

I am truly sorry for the losses you have met with in your attempts to make profit by commerce in this country.

Ingenhousz in his reply of October 24, 1788, says:

This money was invested in the hope of being able to live in the new world on the proceeds of this investment. I am not rich enough to bear such a great loss if Mr. Wharton should prove to be a dishonest man.

It is not certain whether Ingenhousz was on his way to America when he arrived in Paris on the day of the fall of the Bastille (July 14, 1789). Here he received word of the death of his brother which forced him to go to Breda. Being an Austrian citizen he was in danger of being captured by the French revolutionaries and hence he tried to reach England instead of returning to Austria. After a hazardous trip he arrived in London where he learned of the death of Franklin. Owing to the nerve-racking experiences he had gone through on this last trip, his health broke down and he decided to remain in England until peace was restored. Part of the year he stayed in Herforth with his friend Dinsdale and spent the remainder of the time at Bowood, the country residence of the Marquis of Lansdowne (Earl of Shelburne). Here he was offered a room, known as "the laboratory," to carry out some additional experiments on the nutrition of plants, the results of which were published three years before his death in 1799. He was buried in Calne. His wife, to whom he had left his not inconsiderable fortune of 30,000 florins, died the next year.

Ingenhousz was well thought of by his contemporaries. Franklin speaks very highly of him in several letters and Alexander von Humboldt in the introduction to the German edition of his last work states:

Mr. Ingenhousz belongs to the small group of active physicists who possess the fruitful talent of being able to study with admirable energy separate topics and connect them harmonically with phenomena already known.

A bust of Ingenhousz, carrying the inscription "Qua ratione plantae alantur primus perspexit," is found in the arcade of the main building of the University of Vienna, next to those of his illustrious compatriots van Swieten and von Jacquin.

THE SUBJECT VERSUS SPECIAL STUDENT INTERESTS[1]

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MANY of us who are concerned with the teaching of general chemistry are frequently presented with requests—if not demands—by other departments or schools for special courses for different groups. Students in engineering, agriculture, home economics, police administration, hotel administration, nursing, schools or departments as well as premedical and predental students and majors from other science fields come to chemistry departments for a chemistry course that will be of particular use to them in their chosen field. That the title of the first course is usually "general chemistry" is often overlooked.

Many chemistry teachers believe that the most lasting service they can render these students is to give them some understanding of basic principles that are generally applicable to any chemical situation or problem. The principles of chemistry are the same regardless of the field of application, and hence the very logical argument for one introductory chemistry course. That these principles may not be simply applied to complex situations is apparent to anyone with a professional training in chemistry. The novice, the student who takes but one course, and that only because it is required by his chosen curriculum, is soon complaining that he can see no relation between an isolated principle, carefully insulated from undesirable variables, and the more complex problems of his applied field whether it be in the dairy, kitchen, hospital, or elsewhere.

[1] Presented before the Division of Chemical Education at the 120th Meeting of the American Chemical Society in New York, September, 1951.

The student often wants to know what is the best water softener, paint remover, radiator cleaner, soap, cosmetic, depilatory, cigarette, laxative, storage battery, or what have you. This point of view is held despite the fact that conditions of use vary widely and the best answer today is obsolete tomorrow with the development of new products.

Although it is by no means uniform practice, premedical, predental, and engineering students, prospective chemistry majors, and for the most part science majors from other fields respond well to a good general-chemistry course.

There is, however, less satisfaction expressed by majors from the fields of home economics, nursing, agriculture, and others that require some knowledge of biochemistry for which only a very limited time is allowed in the curriculum. It is with this latter group that I am particularly concerned at the moment.

If we attempt to analyze the complaints and look behind the symptoms for the causes, a number of sources of trouble become apparent. For convenience I choose to group these into the following categories: insufficient preparation, student immaturity, improper counseling, insufficient time allotted to attain stated objectives, and failure to use the prescribed chemical foundation in later courses in the applied field.

Insufficient preparation for a general-chemistry course is nothing new. Student difficulties with reading and arithmetic have long been with us. Either these problems are increasing in frequency, or our powers of diagnosis are improving, or both. An increasing number of chemistry departments are using aptitude tests that are essentially arithmetic and reading tests. The data obtained from these tests give direct evidence of too frequent lack of the tools of learning among the students admitted to college. Many schools are operating noncredit clinics where these deficiencies may be removed. The erection of these filters in front of the general-chemistry course, porous as they may be, materially reduces the number of failures.

Student immaturity, as evidenced by lack of purpose and a knowledge of the techniques of effective study, is not peculiar to chemistry students. Since chemistry was moved from the freshman to the sophomore year in the curriculum for most agriculture students at Michigan State College the "gripes" have largely disappeared. These students have had a year's experience with college study and other departments have effected the inevitable and disagreeable weeding process.

Much good work is being done by counseling arrangements of various kinds. Difficulties due to poor counseling can usually be ascribed to a lack of communication between counselor and the teacher who sometimes is forbidden to counsel with his students because he has not been professionally trained in this work. This is particularly acute when it occurs in the sciences where the professional counselor may have no personal or intimate knowledge of the field represented by the students' interests.

The general-education requirements now being imposed upon all students at various institutions, added to the specific requirements frequently laid down by professional societies and accrediting agencies, have developed tremendous pressure on the curriculum of many students. It is no wonder that the word "integration" has skyrocketed in popularity. The increase in our knowledge of chemistry and our desire to tell as complete a story as possible, has collided head on with the expanding list of requirements presented to students. Would it not be preferable to try to teach less and teach it better?

If the department that requires a chemistry course fails to use it later, the students soon realize this and resent the imposition of an apparently unnecessary academic hurdle. This resentment naturally spills over on the chemistry instructor and increases the difficulty of his job.

The argument for the single general-chemistry course implies a reasonable degree of uniformity of student preparation, of intellectual maturity, of purpose, and of the desire for some degree of mastery of the material presented. The extent to which this uniformity is present, of course, depends upon the admission requirements and procedures of the particular institution. Even the casual observer must note the similarity between the changes that took place in secondary education a generation ago and those that are now taking place in the colleges. The increasing percentage of the population that seeks higher education, coupled with the trends in what is called "general education" bring to our doors more students with less uniform preparation and purpose. Despite this increase in numbers, only a minority of high-school graduates go on to college, and college preparation is not the main concern of the secondary teacher.

Reversing the habits of the ostrich, we may hide our heads in our ivory towers and teach only those who come to us on our terms. If this is to be our procedure the work will be taken up by others.

Schools of nursing in colleges and universities are in the process of devising their own "integrated science course" where the student deals with "whole problems." Such chemistry as the professional nurse is to study is covered within the framework of this course. One professed objective of this course is to save time in an overcrowded curriculum and to allow more time for social sciences. As a part of a training program for nurses, this is no concern of the college chemistry teacher but along with this development there is increased emphasis upon the nurse as a professional person, and if the objective is a professional education capped with a baccalaureate degree, the program might be strengthened by the contributions of a competent chemistry teacher.

The State Department of Education is putting pressure on one school of home economics to offer its own introductory chemistry course. This is indicative of a trend which is spreading like a Washington Bureau.

Some of the criticism directed at many conventional chemistry courses are justified. For example, lack of

joint planning with other departments served, duplication of subject matter, and blind adherence to assumed conventions. The remedy, however, would seem to be the correction of faults where they occur and not the complete disregard of what has been learned about chemical education to date. In my opinion it is much easier for a competent teacher to modify his presentation than it is for a novice to teach chemistry. Does it necessarily follow that illustrations of principles taken from the activities of the nurse or homemaker will be any less effective than those from the contact process for the production of sulfuric acid?

If chemistry is to be taught, let it be taught by a good teacher who thoroughly understands his students, their objectives, the objectives of the curriculum in which they are enrolled, and *chemistry!*

Many of our present difficulties would be avoided if we would distinguish clearly between training and education, between some familiarity with a special vocabulary and an understanding of the principles that describe the behavior of all matter. Those who prescribe special curricula must make this choice. Training, if it belongs in a college curriculum, is best performed by the specialist who can mold it to the applied field. Chemical education is best directed by those chemists who are also good teachers.

The chief criticism of textbooks for special courses from the writer's point of view are twofold. They concern accuracy and the mass of material presented. A sampling of these from the shelf of desk copies in almost any chemistry department office will show that all too frequently the presentation is similar to that found in more conventional general-chemistry texts of a generation ago. It is rather obvious that many of these authors while practicing in applied fields have not kept abreast of fundamental chemical theory. One book with a recent date line has the following statement: "According to Raoult's Law, the depression of the vapor pressure is proportional to the number of molecules of a solute dissolved in a given volume of solvent. Thus, by doubling the number of molecules of solute, the vapor pressure of the solvent is doubled." I have heard an instructor insist, over the objections of his students, that bricks are organic in nature because they turn black when a building burns. I have also heard that there is air in only a small part of a lamp bulb since most of it has been pumped out. Errors are no novelty, but if chemists refuse to help special groups they are contributing to the dissemination of misinformation.

The second and more insidious fault is in the bulk of material presented without any attempt at real understanding. The questions at the end of chapters in such books are either strictly factual or require a dogmatic repetition of a statement of principle without adequate explanation being offered, and without real understanding being attained.

Detailed descriptions of the individual metals, their metallurgy, compounds, and uses have largely been abbreviated and in many cases reduced to a general treatment in a chapter or two in the present-day texts, as the inclusion of more theory has required additional space. Is there not much further opportunity for deletion rather than condensation?

I should like to see a text for nonchemistry majors that:

(1) stresses the relationship between structure and properties of substances with suitable illustrations. And incidentally, one that mentions molecules of sodium chloride no more frequently than they are encountered in the laboratory.
(2) gives to the concept of equilibrium an emphasis comparable with its importance in understanding chemical systems.
(3) shows the student something about his environment and frankly does not try to tell him all about it.
(4) will give the student a comprehension of the chemical nature of typical substances and suggest to him where he can find specific answers when needed.
(5) states facts to be learned and ideas to be explored in a measure commensurate with the time available.
(6) will challenge the student to continue on his own without demanding the impossible in a period of 12, 16, 24, or 36 weeks.

There is another point at which just criticism may be leveled at chemistry departments. The proper development of the special course requires more than average insight into the essential nature of our world, its physical structure and composition. This can only be accomplished by marshaling the best brains available. How much of the present difficulty has been brought about by a tendency to shunt off onto less active members of the faculty the responsibility for special courses because the students may be less critical? How often has the treatment of such a course been given a minimum of thought and time because promotion, academic "respectability" and reputation depend upon the list of publications rather than the number of students inspired?

There is no suggestion made here that we compromise in any way the general-chemistry course for students who can use the regular course in general chemistry. The question is what shall we offer, if anything, for those who for one reason or another do not find the time, the inclination, or the opportunity to take the regular course in general chemistry?

It has been pointed out by many others that we are alarmingly short on fundamental research and are not training enough students with the capacity and will to carry out these most important functions. Our strength in the past has been our capacity to adapt the fundamental discoveries from many countries to mass production in our own. The A. C. S. and many other organizations are trying to educate the general public to the need for truly fundamental research and to ob-

tain support for it. Where is a better opportunity to be found than in the classroom, with a large group of students who, while not potential professional scientists, could constitute an informed segment of the population aware of the significance of chemistry in particular and science in general in our civilization?

Some objective thinking and study are needed to determine who should take chemistry and how much. Those making curricula should be encouraged to determine objectively what chemistry has to offer their students and to weigh its advantages against what must be omitted to make room for it.

THE LONG FORM OF THE PERIODIC TABLE

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IT IS a matter of very considerable interest to learn that Professor Coryell (*1*) has proposed independently a long form of the periodic table similar to the one published by us (*2*). In an article on "The periodic table: The *6d-5f* mixed transition groups" (*3*) Coryell mentions a number of interesting points which have to do with the problem of drawing "tie-lines" between elements in the table. In this respect Dr. Coryell differs with us in certain cases. At the outset it should be said that different authors will very likely differ in the way these tie-lines should be drawn. It is a matter of opinion and judgment whether or not a tie-line should be a heavy printed line or a dotted line. As a matter of fact it would be desirable to shade these tie-lines in varying degrees to indicate the different relations that exist between the elements in question. Coryell points out that the limitations of typography restrict the possible shadings of meaning it would be desirable to convey by their use.

It has already been pointed out in the previous publication (*2*) that any form of the periodic table at best will give only an approximate guide to the physical and chemical properties of the elements. Minor differences in the structure of the table will be due to personal preferences of the various investigators. Perhaps the only thing that can be asked of a periodic table is that it would be simple, instructional, and include all the known elements in a single framework.

In these respects there is essentially no difference between our modification and that proposed by Coryell (*1*, *3*). For example, the position of the neutron can be chosen as suggested by Coryell *or* by us. It is located above helium in the latter's table. Its position in a "zeroth" row seems to be a matter of choice.

The positions of thorium, protactinium, and uranium in the periodic table have been debated for a long time. Recent chemical investigations indicate clearly that chemically these elements are homologues of hafnium, tantalum, and tungsten. Thus the tie-lines proposed by Coryell between these elements are definitely justified. However, it is well known that the inclusion of the spectral terms of the ground state of atoms in the periodic table shows that the elements of the same column usually have the same spectral terms.

Although it is not certain that thorium and protactinium have any *5f* electrons, Schuurmans (*4*) found by spectroscopic observation, that uranium has three *5f* electrons in the neutral atom. It is therefore an electronic homologue of neodymium and some recognition of this fact probably should be given in the periodic table.

Another recent publication (*5*) shows that the spectra of europium and americium are quite similar in appearance and that, therefore, it is very probable that americium has seven *5f* electrons. Therefore, although there is very little chemical similarity between the two elements, a tie-line between them is again justified. Perhaps the solid tie-line between tin-hafnium, antimony-tantalum, etc., should be replaced by dotted ones. However this equally well applies to the tie-line between copper-silver-gold or between cesium-francium. In these cases the chemical and physical properties of the elements show much greater variation than, say elements of the halogen group. It may be possible to improve the periodic table by varying the shading of the tie-lines to indicate different degrees of relationship between elements. However, such shading would certainly vary from author to author.

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PROTON WAVES

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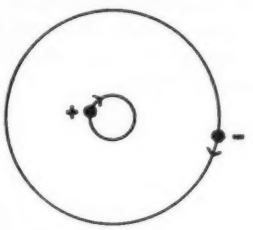
It is the purpose of this note to draw attention to certain features of the wave-mechanical interpretation of atomic structure which appear to have a certain novelty although they are implicit in standard treatments.

In order to prepare the ground the analogues of these features in Bohr's theory of the hydrogen atom will first be discussed, relativity corrections being omitted. As is well known, the energies are correctly deduced if the electronic mass is replaced by the reduced mass and the latter is imagined to be attracted to a fixed center placed at a distance equal to the interparticle distance. This treatment obscures what is actually imagined to happen according to Bohr's theory, since *both* the nucleus and electron rotate about the mass center, as shown in the figure for circular orbits, the two particles preserving the distance between them as if joined by a rigid rod. If m and M are the masses of electron and nucleus, r_1 and r_2 their distances from the mass center and r the interparticle distance,



$$r_1 = \frac{M}{m+M}r \quad r_2 = \frac{m}{m+M}r \quad r = \frac{n^2h^2}{4\pi^2h^2\mu e^2}$$

where μ is the reduced mass and n the principal quantum number. Hence

$$r_1 = \frac{n^2h^2}{4\pi^2me^2} \quad r_2 = \frac{n^2h^2}{4\pi^2Me^2}$$

Since the Bohr radius, $h^2/(4\pi^2me^2)$, is equal to 0.53 Å., it follows that for large values of n, r_2 will be appreciable. Thus if $n = 8$, which corresponds to the experimentally known term value 1714 cm.$^{-1}$, r_2 is approximately 0.02 Å. and, if $n = 50$, r_2 is ± 0.7 Å.

On the emission of light both the electronic and nuclear orbit must simultaneously contract, a rather curious but inescapable result. A similar treatment may be given for elliptical orbits, *both* particles describing such orbits round the mass center, so that the maximum value of r_2 may be appreciable for high eccentricity.

In the calculation of magnetic moments it is worth noting that the replacement of electronic by reduced mass gives incorrect results. This arises from the fact that *both* particles contribute to the orbital magnetic moment, the total moment being $e\omega/2c\ (r_1^2 - r_2^2)$ for circular orbits, where ω is the angular velocity and where the negative sign arises from the fact that electron and nucleus are oppositely charged. The orbital angular momentum is

$$mr_1^2\omega + Mr_2^2\omega = \mu r^2\omega = \frac{nh}{2\pi}$$

Hence the ratio

$$\frac{\text{angular momentum}}{\text{orbital magnetic moment}} = \frac{2\mu cr^2}{e(r_1^2 - r_2^2)} = \frac{2\mu c(m+M)^2}{(M^2 - m^2)e} = \frac{2mMc}{e(M-m)}$$

This ratio is usually quoted as $2mc/e$ and if m were replaced by μ we should obtain an incorrect result. Likewise the orbital magnetic moment is

$$\frac{nhe}{4\pi mc}\frac{(M-m)}{M}$$

instead of the usually quoted value,

$$\frac{nhe}{4\pi mc}$$

It follows that the ratio of spin magnetic moment to the orbital magnetic moment is not 2, but $2M/(M-m)$.

In a similar way the diamagnetic susceptibility due to an electron and a proton, both of which describe circular orbits, is not $-e^2r^2/(4mc^2)$, as applies to an electron alone, but

$$\frac{-e^2}{4c^2}\left(\frac{r_1^2}{m} + \frac{r_2^2}{M}\right) = \frac{-e^2}{4c^2}\frac{(M^3+m^3)}{mM(m+M)}r^2$$

Although these corrections are small they could give rise to observable effects and they bring out the advantage of treating the two particles in an equivalent fashion, a principle which it is the purpose of this note to emphasize, and which may be carried over to the wave-mechanical treatment. The wave equation for a hydrogen atom may be separated into an equation describing the motion of the mass center and one involving the reduced mass. As an example of the solution of the latter equation we may quote

$$\psi\ (\text{ground state}) = \left(\frac{1}{\pi a_0}\right)^{3/2} e^{-r/a_0}$$

where $a_0 = h^2/(4\pi^2\mu e^2)$; $\psi\psi^* d\tau$ gives the probability of finding an electron in an element of volume $d\tau$ distant r from the nucleus where ψ^* is the complex conjugated ψ. All this is of course correct, but the usual treatment then goes on to consider electron distributions round a fixed nucleus, instead of treating the two particles in an equivalent manner. This may be done for a fixed mass center by the use of the consideration that $\psi\psi^* d\tau$ gives the probability of finding an electron in an element of volume $d\tau$ distant r_1 from the mass center and *at the same time* the probability of finding the nucleus in the element of volume $d\tau$ distant r_2 from the mass center, the two particles and the mass center lying on

the same line. If this were not so, the mass center would not remain fixed. The wave functions describing the behavior of the electron and the nucleus are then obtained by replacing r by the equal values

$$\frac{m + M}{M} r_1$$

and

$$\frac{m + M}{m} r_2$$

giving

$$\psi \text{ (electron)} = \left(\frac{1}{\pi a_0}\right)^{3/2} e^{-\frac{(M+m)}{M}\frac{r_1}{a_0}}$$

and

$$\psi \text{ (nucleus)} = \left(\frac{1}{\pi a_0}\right)^{3/2} e^{-\frac{(m+M)}{m}\frac{r_2}{a_0}}$$

These two functions are equal when r_1 and r_2 are given their appropriate values, but they may be considered separately as giving the ψ values as functions of all values of r_1 and r_2, and they are of course not equal at equal distances from the mass center. The nuclear function decays much more rapidly than the electronic. The nuclear "cloud pattern" is thus a kind of shrunken version of that due to the electron. Similar results are obtained for excited states, for which the wave functions contain as factors polynomials in r.

We thus have the result that the nuclear cloud is spread over the whole atom, just as is the electron cloud, and at first sight the interpenetration of these two clouds seems unlikely. Such a view overlooks the statistical nature of the cloud patterns, and it is clear that the proton and electron always lie on opposite sides of the mass center. We thus have the idea of proton waves permeating the atom, and it follows that any discussion of the combination of two hydrogen atoms in terms of atomic electron-wave functions should also allow for the proton waves. This is not likely to lead to any serious correction to energy values. The way is opened up, however, for the concept of proton waves spreading throughout a molecule, albeit with very great attenuation, and we have here a possible factor in prototropic changes in chemistry. In principle the effects will apply to other nuclei but will be much smaller than for hydrogen.

QUALITATIVE ADEQUACY OF PHLOGISTON

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THE usual treatment of phlogiston in freshman chemistry texts gives the impression that it was an absurd notion and that people were rather foolish to hold it. Admittedly it was a false system, was not quantitative, and could not become so without increasing its complications. But it did predict qualitative results rather well and it was used by such competent persons as Scheele, Priestley, and Cavendish in organizing their thought and in planning their work.

Since phlogistonism never became a quantitative system, the following account will not attempt to adjust itself to the quantities of material used and produced. However, equations will be written to make them agree with the Law of Conservation of mass, which was tactily in the minds of many eighteenth century chemists. Also, since the whole procedure is very anachronistic it is adapted to a number of reactions which perhaps were unknown to Scheele in 1774 when he wrote his text on "Fire and Air." They are: (1) oxygen and hydrogen form water, (2) metal oxides and hydrogen form water and the metal, and (3) steam and carbon form hydrogen and carbon monoxide. Otherwise, it generally follows Scheele in his theory of fire and air. Heat is represented as the union of oxygen and phlogiston, one of the basic points in his theory.

Modern chemists are accustomed to discussions illustrated by equations based on formulas which show the constitution of substances. This is perhaps one of the reasons why it is so difficult to read the papers of the phlogistonists. Their names for the various substances are unfamiliar or misleading and there is very little effort to formulate reactions by equations. The author has felt that an adaption of the modern Berzelian symbolism to the phlogiston theory would make it clearer to students of elementary chemistry. The following symbolism has been devised. Phlogiston is represented by the Greek letter *phi*, Φ. Oxygen is Fa from Scheele's name *fire air*. Heat, being the compound of fire air and phlogiston, is FaΦ. Water is W and hydrogen, WΦ. For metals, the calxes are represented by the modern

symbols with an X placed before them. Thus zinc oxide is XZn and zinc is XZnΦ. Iron was known to Scheele in both ferrous and ferric states. XFe will have to be ferric oxide; XFeΦ, ferrous oxide; and XFeΦ$_2$ iron metal. The same system could be used for any metal of multiple valence, but it has seemed to the author that one metal was enough to illustrate this possibility.

The acids and the nonmetals present a greater complication. Scheele knew that nitric oxide, nitrous acid, nitrogen dioxide, and nitric acid stood in that order. He also knew that ammonia was converted to an inert gas by oxidizing agents and Cavendish connected nitrogen to nitric acid. So, based on the old name *aqua fortis*, the following formulas result:

Nitric acid	Af	Nitric oxide	AfΦ$_2$
Nitrogen dioxide	AfΦ	Nitrogen	AfΦ$_4$
Nitrous acid	AfΦ$_2$	Ammonia	AfΦ$_3$

Similarly, sulfuric acid (*oil of vitriol*) and carbonic acid (*aerial acid*) and their reduction products are

Sulfuric acid	Ov	Carbon dioxide	Aa
Sulfur dioxide	OvΦ	Carbon monoxide	AaΦ
Sulfur	OvΦ$_2$	Carbon	AaΦ$_2$
Hydrogen sulfide	OvΦ$_3$		

Chlorine has its modern symbol and hydrochloric acid is ClΦ.

One further explanation is perhaps necessary. The hydrogen theory of acids is far in the future. It was known that acid anhydrides held water firmly and that quick lime took on water to form slaked lime. These were looked on as compounds in which the water continued to exist independently. In other words, water was bound as a water of hydration and was neglected unless its presence altered the results obtained.

Salts were a combination of a calx and an acid, thus:

Zinc sulfate	XZn.Ov	Mercuric chloride	XHg.ClΦ
Zinc nitrate	XZn.Af	Ferrous chloride	XFeΦ.ClΦ
Zinc sulfite	XZn.OvΦ	Ferric chloride	XFe.2ClΦ
Zinc nitrite	XZn.AfΦ$_2$		

Ferric chloride is formulated XFe.2ClΦ because of the following reactions:

$$\underset{iron}{XFe\Phi_2} + \underset{chlorine}{Cl} \rightarrow \underset{ferrous\ chloride}{XFe\Phi.Cl\Phi}$$

$$\underset{ferrous\ chloride}{XFe\Phi.Cl\Phi} + \underset{chlorine}{Cl} \rightarrow \underset{ferric\ chloride}{XFe.2Cl\Phi}$$

Based on this symbolism one can write the following equations:

Metal oxides and salts react with acids

$$\underset{calcium\ oxide}{XCa} + \underset{sulfuric\ acid}{Ov} \rightarrow \underset{calcium\ sulfate}{XCa.Ov}$$

$$\underset{Sodium\ chloride}{XNa.Cl\Phi} + \underset{sulfuric\ acid}{Ov} \rightarrow \underset{sodium\ sulfate}{XNa.Ov} + \underset{hydrochloric\ acid}{Cl\Phi}$$

Metals with dilute acids

$$\underset{iron}{XFe\Phi_2} + \underset{water}{W} + \underset{sulfuric\ acid}{Ov} \rightarrow \underset{hydrogen}{W\Phi} + \underset{ferrous\ sulfate}{XFe\Phi.Ov}$$

Note that the dilute acid furnishes the water for the hydrogen. If water is absent, one gets such reactions as

$$\underset{zinc}{XZn\Phi} + \underset{sulfuric\ acid}{2Ov} \rightarrow \underset{zinc\ sulfate}{XZn.Ov} + \underset{sulfur\ dioxide}{Ov\Phi}$$

$$\underset{silver}{XAg\Phi} + \underset{nitric\ acid}{2Af} \rightarrow \underset{silver\ nitrate}{XAg.Af} + \underset{nitrogen\ dioxide}{Af\Phi}$$

$$\underset{copper}{3XCu\Phi} + \underset{nitric\ acid}{4Af} \rightarrow \underset{copper\ nitrate}{3(XCu.Af)} + \underset{nitric\ oxide}{Af\Phi_2}$$

Metals with salts

$$\underset{zinc}{XZn\Phi} + \underset{copper\ sulfate}{XCu.Ov} \rightarrow \underset{copper}{XCu\Phi} + \underset{zinc\ sulfate}{ZXn.Ov}$$

$$\underset{copper}{XCu\Phi} + \underset{silver\ nitrate}{XAg.Af} \rightarrow \underset{silver}{XAg\Phi} + \underset{copper\ nitrate}{XCu.Af}$$

Scheele had previously studied arsenic trioxide and manganese dioxide. In these investigations he had observed that increasing amounts of phlogiston converted an acidic substance into one which had basic properties. He had prepared potassium manganate and had seen the color of permanganate in solution. He had observed that manganese dioxide would not give colorless solutions in acids unless it was reduced to the manganous state. Using Pm for permanganic acid, the following formulas and equations result:

Permanganic acid	Pm	Manganese dioxide	PmΦ$_2$
Manganic acid	PmΦ	Manganous oxide	PmΦ$_2$

$$\underset{manganese\ dioxide}{Pm\Phi_2} + \underset{nitric\ acid}{Af} \rightarrow \text{no reaction}$$

$$\underset{manganese\ dioxide}{Pm\Phi_2} + \underset{nitrogen\ dioxide}{Af\Phi} \rightarrow \underset{manganous\ nitrate}{Pm\Phi_2.Af}$$

$$\underset{manganese\ dioxide}{Pm\Phi_2} + \underset{sulfur\ dioxide}{Ov\Phi} \rightarrow \underset{manganous\ sulfate}{Pm\Phi_2.Ov}$$

$$\underset{manganese\ dioxide}{Pm\Phi_2} + \underset{iron}{XFe\Phi_2} + \underset{sulfuric\ acid}{2Ov} \rightarrow \underset{manganous\ sulfate}{Pm\Phi_2Ov} + \underset{ferrous\ sulfate}{XFe\Phi\cdot Ov}$$

$$\underset{manganese\ dioxide}{2Pm\Phi_2} + \underset{hydrochloric\ acid}{4(Cl\Phi)} \rightarrow$$

$$\underset{manganic\ tetrachloride\ (brown)}{Pm\Phi_2.2(Cl\Phi)} \rightarrow \underset{manganous\ chloride\ (colorless)}{Pm\Phi_2.Cl\Phi} + \underset{chlorine}{Cl}$$

Scheele discovered chlorine by this last reaction and described its properties and its reactions with metals. In all of these reactions manganese dioxide is represented as gaining phlogiston from some substance. Scheele noted that when manganese dioxide and sulfuric acid were heated to a high temperature the manganese dissolved, forming manganous sulfate.

$$\underset{manganese\ dioxide}{Pm\Phi_2} + \underset{sulfuric\ acid}{Ov} + \underset{heat}{Fa\Phi} \rightarrow \underset{manganous\ sulfate}{Pm\Phi_2.Ov} + \underset{oxygen}{Fa}$$

Where did the necessary phlogiston come from? He believed that it came from the heat and that the same reaction was not possible with more volatile acids since they were boiled off before the proper concentration of heat was reached. Further investigation showed that the other product of this reaction was a gas which was a notable supporter of combustion. This gas he called

fire air and it is symbolized in the equation with Fa. Note that fire air is oxygen and not phlogiston. He went on to heat other substances which he knew could be reduced and obtained the same gas from these reactions also. The calxes of gold and silver behaved similarly to that of mercury.

$$\underset{\substack{\text{calx of}\\\text{mercury}}}{\text{XHg}} + \underset{\text{heat}}{\text{Fa}\Phi} \rightarrow \underset{\text{mercury}}{\text{XHg}\Phi} + \underset{\text{oxygen}}{\text{Fa}}$$

$$\underset{\text{nitric acid}}{\text{Af}} + \underset{\text{heat}}{\text{Fa}\Phi} \rightarrow \underset{\text{nitrogen dioxide}}{\text{Af}\Phi} + \underset{\text{oxygen}}{\text{Fa}}$$

$$\underset{\substack{\text{potassium}\\\text{nitrate}}}{\text{XK.Af}} + \underset{\text{heat}}{2(\text{Fa}\Phi)} \rightarrow \underset{\substack{\text{potassium}\\\text{nitrite}}}{\text{XK.Af}\Phi_2} + \underset{\text{oxygen}}{2\text{Fa}}$$

This final reaction, even though the temperature caused the glass retort to soften, was described by Scheele as the cheapest way of preparing oxygen.

Scheele believed that heat was a substance, and the passage of the subtle particles of heat through the solid glass was only a marvel and not an objection to the theory. Light was a higher state of phlogistication and he showed that it acted on silver chloride suspensions to form silver and hydrochloric acid. Heat alone would not accomplish this.

$$\underset{\substack{\text{silver}\\\text{chloride}}}{\text{XAg.Cl}\Phi} + \underset{\text{light}}{\text{Fa}\Phi_2} \rightarrow \underset{\text{silver}}{\text{XAg}\Phi} + \underset{\substack{\text{hydrochloric}\\\text{acid}}}{\text{Cl}\Phi} + \underset{\text{heat}}{\text{Fa}\Phi}$$

Scheele went on to demonstrate the composition of heat synthetically to accompany his analytical demonstrations. Fire air should remove phlogiston from substances and form heat. He tried a large number of substances, but a few will suffice to show the nature of his thought.

$$\underset{\text{sulfur}}{\text{Ov}\Phi_2} + \underset{\text{oxygen}}{\text{Fa}} \rightarrow \underset{\text{sulfur dioxide}}{\text{Ov}\Phi} + \underset{\text{heat}}{\text{Fa}\Phi}$$

$$\underset{\text{sulfur}}{\text{Ov}\Phi_2} + \underset{\text{oxygen}}{2\text{Fa}} \rightarrow \underset{\text{sulfuric acid}}{\text{Ov}} + \underset{\text{heat}}{2(\text{Fa}\Phi)}$$

$$\underset{\text{copper}}{\text{XCu}\Phi} + \underset{\text{oxygen}}{\text{Fa}} \rightarrow \underset{\text{copper oxide}}{\text{XCu}} + \underset{\text{heat}}{\text{Fa}\Phi}$$

$$\underset{\substack{\text{calcium polysulfide}\\\text{("liver of lime")}}}{\text{XCa.Ov}\Phi_2\text{.Ov}\Phi_2} + \underset{\text{oxygen}}{3\text{Fa}} \rightarrow \underset{\substack{\text{calcium}\\\text{sulfate}}}{\text{XCa.Ov}} + \underset{\text{sulfur}}{\text{Ov}\Phi_2} + \underset{\text{heat}}{3(\text{Fa}\Phi)}$$

$$\underset{\text{iron}}{\text{XFe}\Phi_2} + \underset{\text{oxygen}}{2\text{Fa}} + \underset{\text{water}}{\text{W}} \rightarrow \underset{\substack{\text{ferric}\\\text{hydroxide}}}{\text{XFeW}} + \underset{\text{heat}}{2(\text{Fa}\Phi)}$$

The last reaction went especially well when the iron filings were moistened with water. So the reaction between the calx of iron and water merits notice in the equation. The last two reactions were very slow and did not become hot to touch, but Scheele demonstrated by the use of a thermometer that heat was liberated in them.

Scheele considered two other interpretations of these reactions. One was held by Priestley and many other phlogistonists—namely that the phlogiston lost in burning was taken up by the air, causing it to "mat down," thus explaining the decrease in volume when things are burned in an enclosed sample of air. Scheele did not adopt this concept because he found that the residual gas (nitrogen) was lighter than air and that fire air was heavier. He reasoned that if air absorbed phlogiston and contracted it would have to become heavier, not lighter. The other possibility—that the air lost in burning was taken up by the substance burned—he discarded because he could not recover any gas from the potassium sulfate obtained when potassium sulfite absorbed oxygen, this reaction he formulated thus:

$$\underset{\substack{\text{potassium}\\\text{sulfite}}}{\text{XK.Ov}\Phi} + \underset{\text{oxygen}}{\text{Fa}} \rightarrow \underset{\substack{\text{potassium}\\\text{sulfate}}}{\text{XK.Ov}} + \underset{\text{heat}}{\text{Fa}\Phi}$$

The easiest way to represent the burning of hydrogen would be

$$\underset{\text{hydrogen}}{\text{W}\Phi} + \underset{\text{fire air}}{\text{Fa}} \rightarrow \underset{\text{water}}{\text{W}} + \underset{\text{heat}}{\text{Fa}\Phi}$$

This formulation of hydrogen as water and phlogiston is in accord with the formation of water and a metal in the reduction of metal oxides and with the formation of water gas from steam and carbon.

$$\underset{\text{copper oxide}}{\text{XCu}} + \underset{\text{hydrogen}}{\text{W}\Phi} \rightarrow \underset{\text{copper}}{\text{XCu}\Phi} + \underset{\text{water}}{\text{W}}$$

$$\underset{\text{carbon}}{\text{Aa}\Phi_2} + \underset{\text{steam}}{\text{W}} \rightarrow \underset{\text{carbon monoxide}}{\text{Aa}\Phi} + \underset{\text{hydrogen}}{\text{W}\Phi}$$

Scheele himself did not accept this solution, since the density of hydrogen compared to that of oxygen made it seem absurd that the weight of tangible product, W, was carried by the hydrogen. His proposal was that oxygen consisted of water, phlogiston, and some "saline" substance. Hydrogen was made of the same three ingredients in different proportions. Only the water had any great weight. The saline principle and the phlogiston were so light that the mass escaping as heat was indetectable. Thus the reaction would look something like

$$\underset{\text{oxygen}}{\text{SW}_8\Phi_2} + \underset{\text{hydrogen}}{\text{SW}\Phi_6} \rightarrow \underset{\text{water}}{9\text{W}} + \underset{\text{heat}}{2(\text{S}\Phi_4)}$$

After Lavoisier proved that the weight of mercuric oxide equaled that of the mercury and oxygen formed in its decomposition, Scheele was forced into some such formulations as

$$\underset{\text{mercuric oxide}}{\text{XHgW}_8} + \underset{\text{heat}}{\text{S}\Phi_4} \rightarrow \underset{\text{mercury}}{\text{XHg}\Phi_2} + \underset{\text{oxygen}}{\text{SW}_8\Phi_2}$$

$$\underset{\text{mercuric oxide}}{\text{XHgW}_8} + \underset{\text{carbon}}{\text{Aa}\Phi_2} \rightarrow \underset{\text{mercury}}{\text{XHg}\Phi_2} + \underset{\text{carbon dioxide}}{\text{AaW}_8}$$

Again in the first reaction the heat was assumed to have an indetectably small mass. Scheele reports some experiments which demonstrate the wonderful fact that the amounts of water in some substances are just enough to supply those in the products obtained. This was the end of phlogiston. The mere fact that the calx was heavier than the metal did not cause its abandonment. The phlogistonists had known that fact for almost a century. There were several explanations advanced, the commonest being that the calx gathered to itself heat, water, or some gas after it had lost phlogiston. Lavoisier was trying to determine what this substance was when he started his experiments. After the discovery of oxygen, Lavoisier noted that the masses

of all substances, both products and reactants, could be accounted for on the basis of oxygen-gain in burning. No simple phlogistic system could do this.

LITERATURE

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A SIMPLE PROCEDURE FOR PREPARING PINACOL HYDRATE

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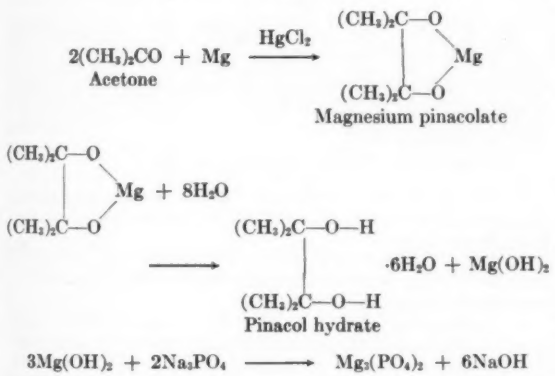
HOLLEMAN'S[1] method for preparing pinacol hydrate by the reduction of dry acetone with magnesium, activated with mercuric chloride, in dry benzene appears in "Organic Syntheses."[2] An adaptation of this method also appears in Fieser's laboratory manual.[3] We have found the following to be a shorter and more simple procedure for this method.



$$3Mg(OH)_2 + 2Na_3PO_4 \longrightarrow Mg_3(PO_4)_2 + 6NaOH$$

Three grams of magnesium turnings were placed in a 200-ml. round-bottomed ring-necked flask and 0.5 g. of mercuric chloride was added. To this was added 10 ml. of acetone from a dry graduated cylinder. The flask was shaken with a rotary motion to give the contents a swirling movement until the reaction started. It was then quickly connected to a water-cooled reflux condenser and 40 ml. more of acetone was added in small portions with shaking.

After the reaction had run for 15 to 20 minutes, another portion of 25 ml. of acetone was added and the flask shaken periodically to avoid the formation of a cake inside the flask. At the end of 30 to 45 minutes, heating over a steam bath was necessary to maintain a rapid reflux rate.

After one hour from the start the reaction had practically stopped and the flask was disconnected from the reflux condenser and connected by means of a one-hole rubber stopper and glass tube to a vacuum line (aspirator). While continuing to heat the flask on the steam bath, a vacuum was applied to remove the unused acetone. By rotating the flask and heating simultaneously the contents of the flask (magnesium pinacolate, etc.) became powder-dry in 5 or 10 minutes.

At this point the reflux condenser was reconnected to the flask and 100 ml. of water containing 5 g. of trisodium phosphate (technical grade) was added through the reflux condenser. The contents of the flask warmed up somewhat on the addition of the water. The mixture was refluxed over a low flame and wire gauze for 15 minutes. The hot mixture was then filtered using a small vacuum filter assembly (60-mm. Büchner funnel and 250-ml. filter flask). The filtrate was chilled in an ice-water bath, stirred vigorously, and seeded. When crystallization seemed to be complete the cold mixture was filtered, using the small vacuum filter assembly again. The crystals were washed in the funnel with 25 ml. of ice water. The vacuum was applied until the crystals were reasonably dry (15 to 20 minutes). They were then washed with 25 ml. of petroleum ether and allowed to dry once again.

It was found that a satisfactory yield (6 to 8 g.) of pinacol hydrate of high purity could be prepared within two hours, using a minimum amount of equipment and ordinary and inexpensive reagents. It will be observed that the use of benzene has been eliminated entirely from the preparation, the relative amount of mercuric chloride has been greatly reduced, and ordinary acetone has replaced the dried product. Trisodium phosphate was added to improve the filtering characteristics of the magnesium residue.

[1] HOLLEMAN, F., *Rec. trav. chim.*, **25**, 206 (1906).
[2] "Organic Syntheses," John Wiley & Sons, Inc., New York, 1932, Coll. Vol. I, p. 448.
[3] FIESER, L. F., "Experiments in Organic Chemistry," 2nd ed., p. 97.

THE THERMOMATIC STEAMER

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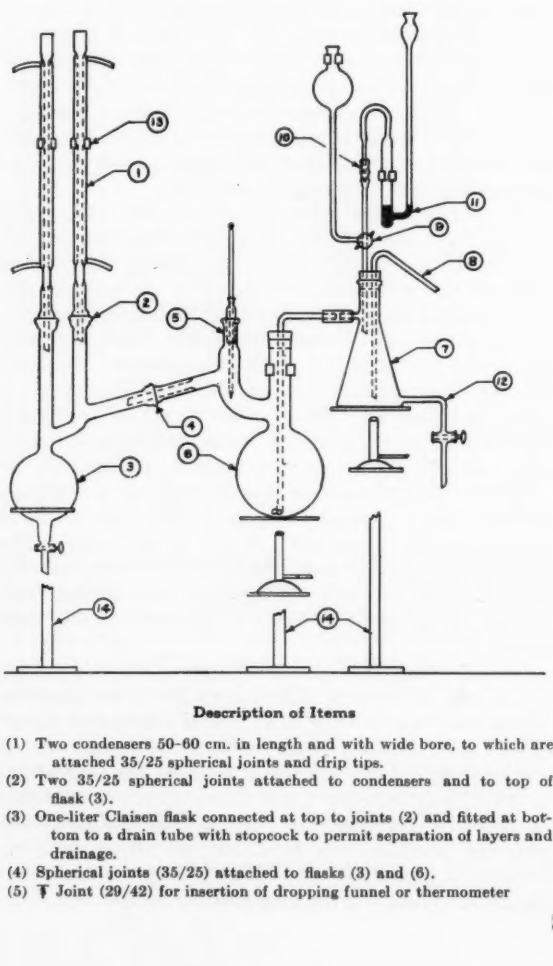
WITHOUT being entirely aware of it, the beginning student in organic chemistry carries out a variety of processes for isolation and purification of compounds known as *unit operations*. Among these are extraction, filtration, crystallization, distillation (atmospheric, vacuum, steam), etc. Facing up to the problem of basic improvements in these processes I reached the conclusion that of all these unit operations, steam distillation carried out by conventional methods is often the messiest and least satisfying to the student, and that something should be done about it. Hence the construction of what I choose to call the unitized "thermomatic steamer," an assembly of apparatus intended to be kept together, ready for immediate use and easy to clean without tearing down, to the end that the next operator can take over without interruption.



Description of Items

(1) Two condensers 50–60 cm. in length and with wide bore, to which are attached 35/25 spherical joints and drip tips.
(2) Two 35/25 spherical joints attached to condensers and to top of flask (3).
(3) One-liter Claisen flask connected at top to joints (2) and fitted at bottom to a drain tube with stopcock to permit separation of layers and drainage.
(4) Spherical joints (35/25) attached to flasks (3) and (6).
(5) ₮ Joint (29/42) for insertion of dropping funnel or thermometer
(6) Two-liter Claisen flask to which is sealed the joint (4) and a 13-mm. delivery tube for passage of distillate.
(7) The boiler: 1-liter suction flask with side arm attached to Claisen steam inlet.
(8) 8-mm. inlet tube for steam or water.
(9) 3-way stopcock with 4-mm. bore.
(10) ₮ joint (10/30) for attachment of safety tube (11).
(11) Safety tube partly filled with mercury, bromobenzene, or other dense liquid to allow release of excessive pressure from the boiler.
(12) Drain tube and stopcock to drain boiler contents.
(13) Clamps for attachment of rods No. (14). The numerous clamps and rings necessary for mounting the apparatus are not all shown, but their positions are obvious.
(14) Ring stand rods (3 ft. × ½ in.) to which the pieces are firmly fastened. These are preferably Fisher Flexaframe rods fastened to a plywood baseboard. The vertical rods should be tied together by a horizontal rod to prevent lateral distortion.

Not numbered: steam inlet to Claisen flask (6). Note that this has two delivery tips at the bottom and these are turned outwards at right angles. This feature prevents blasting of steam at the bottom of (6) and eliminates bumping. With two openings it has not yet been observed that plugging has ever occurred despite accumulation of much precipitated material during some distillations.

Caution: Be sure to anneal the various glass pieces thoroughly after fabrication, particularly flask (7), this being subjected to intense heating which puts a great strain on the drain tube seal.

Over-all dimensions of the assembly: Width, 32 in., Height, 48 in., Depth, 10 in.

The construction and assembly of the apparatus (see the figure) should be obvious, the only problem being a certain amount of modification on standard pieces of glass equipment which are not difficult for the average glass blower.

The essential stages of operation are:

1. Charge the two-liter Claisen flask (6) by attaching a dropping funnel (not shown) at point (5) and running in the mixture to be steam-distilled.

2. Introduce steam or hot water into the boiler (7) through the tube (8). Light both burners and collect distillate in flask (3). Periodically drain (3).

3. Adjust 3-way stopcock (9) so that a slow trickle of water enters the boiler from the spherical flask to replace the steam lost.[1] This insures uninterrupted operation until all the distillables have been removed from flask (6). When not adding water connect boiler to safety tube (11).

4. To stop the distillation and empty flask (6) turn stopcock (9) so that the boiler is closed off from the atmosphere and extinguish the burners; to hasten cooling of flask (7) drain water to a low level then close stopcock again; condensation of steam in (7) sucks back the mixture from (6).

5. Drain (7) again, recharge both flasks, and proceed with the next distillation.

It is obvious that the direct connection of Claisen flask (6) to (3) without an intervening condenser permits the distillation of high-melting materials that might solidify in conventional apparatus—a considerable advantage in this construction. Also, much cooling

[1] Addition of water must be at such a rate that the steam pressure in (7) does not drop; otherwise the contents of (6) will suck back into (7).

and condensation take place in (3) itself so that the condensers above are not so likely to be overloaded. It should be noted that the inclusion of the drip tips in the condensers and the flask (6) protects the joints (2) and (4) from the most direct action of the vaporous products; thus the lubricant used in the joints is better preserved.

The boiler (7) when heated with a large Meker or Fisher burner delivers an abundant supply of steam to flask (6) without any bumping. In fact, operation by this method is much smoother than by use of steam taken from the house line, which frequently produces excessive agitation in the distilling flask due to pressure surges.

The condensers used should be no smaller than those shown. Two of them are certainly necessary. Because of the high latent heat of steam and the volume passed during distillation any change in condensers should be in the direction of larger ones, preferably of the bulb type.

THE 4 IN 1 CHEM-MIXER

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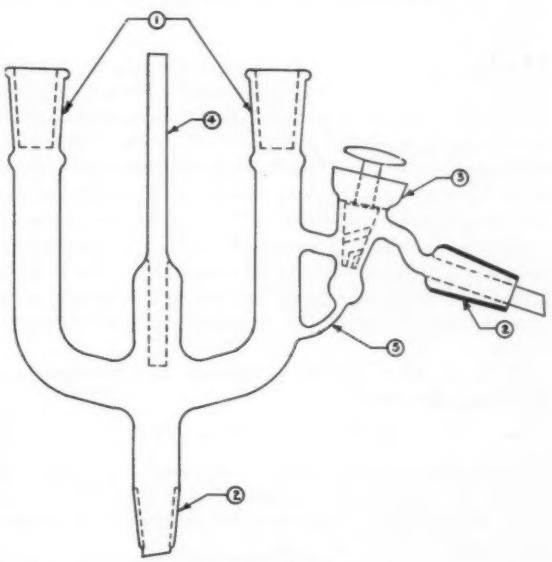
THE three-neck flask is a very valuable unit in the apparatus kit of the organic chemist. It is however not a low cost item when provided with the desired standard taper joints[1] and if a half dozen or so are necessary for a variety of processes, the investment is considerable. For insertion of a proper stirring device into such a flask the neck size should be 29/42 or larger; such necks make it inconvenient to operate with a flask any smaller than 500 ml. in volume. As an added consideration one realizes that whenever a series of operations can be carried out in one vessel without dismantling and emptying, the experimenter can save much valuable time.

The result of these deliberations led to the design and construction of the apparatus shown which I choose to call the "4 in 1 Chem-Mixer." Attached to the figure is a description of items involved in its construction.

By use of the "Chem-Mixer" one may rely principally on one-necked flasks of any size all of which can be adapted to fit the new device. The unit is most conveniently employed when fastened in place with the customary stirrer, dropping funnel, and reflux condenser likewise in position. Thereupon the side tube (2) is atttached to a takeoff and condenser, but the large stopcock (3) is closed for reflux operation. After the main reaction has taken place the solvent or volatile products are removed by distillation with the stopcock in the open position. After these several operations (four in number: stirring, addition, reflux, and distillation) the reaction flask and its contents are removed from the unit, which is then ready to accommodate another batch of reactants without interruption.

It should be noted that the vertical bearing (4) for the stirring shaft can be easily adjusted for inclusion of a mercury seal, or, by a simpler method, attachment of soft rubber tubing over the top, which in many cases gives a perfectly acceptable seal.



Description of Items

(1) 29/42 standard-taper joints into which are fitted a dropping funnel and reflux condenser. These joints are used full length as obtained from supplier and curved at the bottom for attachment to the center tube which connects to the reaction vessel.

(2) 29/42 standard-taper male joints for connection to reaction vessel and to collection unit for distillation.

(3) Vacuum-type stopcock with three holes as shown. When turned 180° from the position in the figure any condensate entering the stopcock body will drain back into the flask through tube (5).

(4) 5–7-mm. tube which acts as bearing for stirring shaft.

(5) Drain tube attached to bottom of stopcock (3). This drains fluid collecting in stopcock. It also serves the important purpose of strengthening the side arm attachment against accidental breakage.

[1] Currently a 500-ml. 3-neck flask with 29/42 joints cost $7.10. A 1000-cc. flask of same type $9.00.

FRAGMENTS OF CHEMISTRY

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TIME was, before the Hatchet Men—commonly known as "Referees"—had usurped editorial functions, when an author could introduce into his paper many interesting and useful bits of information even though they had little bearing on the main theme of the paper. Thus, much useful information was conserved to be used later by other investigators while, at the same time, the paper was usually enlivened thereby. Indeed, such information often stimulated others to undertake researches which, otherwise, they would not have begun. Today, all this has been changed; everything not having a direct bearing on the problem in hand is eliminated. The final paper gives one the impression that nothing has been left undone, and that there is nothing more that need be done. Such papers are not very inspiring. It would seem that all the red meat has been trimmed off the literary roast and all that remains is bone and fat and gristle. And, when a paper is published, the introduction and discussion are printed in large type while the data, if appearing at all, and pertinent experimental details appear in type so small that it requires microfilm equipment to read it. Is this not a case of the horse pushing the cart? The most important element in a paper is the result presented in the form of data along with an evaluation of their precision.

On the basis of considerations such as these, this author has been led to publish some fragments of chemistry, chiefly in the way of observations which have been accumulated over the years and which have been touched upon only briefly in earlier papers or, otherwise, have been left unreported. The brief sketches which follow are, for the most part, undocumented. On request, the author would be glad to supply such further information as he may have.

1. REACTION OF POTASSIUM AMIDE WITH ALIPHATIC POLYHALIDES IN LIQUID AMMONIA[1]

POTASSIUM amide is a moderately strong electrolyte and is very soluble in liquid ammonia. It is an ammono base and, in liquid ammonia, reacts in a manner similar to that of potassium hydroxide in water. However, the amide ion is much less strongly electronegative than is the hydroxyl ion and, therefore, we frequently find it acting as a reducing agent. Thus, potassium amide is oxidized to potassium hydroxide and potassium nitrite by molecular oxygen in liquid ammonia. The solid is likewise oxidized by oxygen but, unless the oxygen pressure is kept very low, the reaction takes place explosively.

With an alkyl halide, potassium amide reacts according to the equation:

$$KNH_2 + RX = RNH_2 + KX \qquad (1)$$

The reaction is quantitative and rapid. Actually, of course, it is the amide ion that reacts; however, for the sake of simplicity, equations will be written in molecular form.

The reactions of methylene halides seem not to have been investigated. Chloroform and potassium amide react in the sense of Franklin's "ammonia system of chemistry" potassium cyanide being formed quantitatively. The reaction may be formulated somewhat as follows:

$$HCCl_3 \xrightarrow{+KNH_2} HC(NH_2)_3 \xrightarrow{-NH_3} HCN \xrightarrow{+KNH_2} KCN \qquad (2)$$

When we come to carbon tetrachloride and hexachloroethane, complex and deep-seated reactions take place which cannot be described in terms of Franklin's ammonia system which involves substitution of NH_2 for chlorine and subsequent deammoniation of the product.

(1) *Carbon Tetrachloride.* According to Franklin's system, we should expect a salt of cyanamide $NCNH_2$, or some related substance, as end product of reaction. Actually, the observed results of reaction are as follows:

(*a*) 90 per cent of the carbon appears as potassium cyanide.

(*b*) Nitrogen is evolved, the ratio of nitrogen atoms to mols of cyanide being in the proportion of 2:3.

(*c*) Hydrogen is evolved along with the nitrogen in the proportion of one atom of hydrogen to three atoms of nitrogen.

(*d*) The ratio of chloride to cyanide is 4.5:1.

(*e*) The ratio of mols of amide to mols of tetrachloride reacted is slightly greater than five (5.25).

The principal product of reaction is potassium cyanide. The CN/N ratio of 3:2 is doubtless significant. The principal reaction might be written:

$$3CCl_4 + 15KNH_2 = 12KCl + 3KCN + N_2 + 10NH_3 \qquad (3)$$

The nitrogen must be formed at some stage of the deammoniation process; potassium cyanamide is a stable, difficultly soluble compound. In all likelihood, the hy-

[1] The observations which follow were accumulated by Dr. Norman L. Cox in the Brown Laboratories during the years 1931–33; they have not been reported heretofore.

drogen is formed in reactions that involve the 10 per cent of carbon that does not form cyanide. The excess amide is probably involved in these side reactions.

(2) *Hexachloroethane.* If hexachloroethane reacted according to Franklin's ammonia system, we should expect cyanogen as end product. Actually, a deep-seated reaction takes place which results in the complete destruction of the hexachloride molecule.

Although difficultly soluble in liquid ammonia, the hexachloride when finely divided, reacts readily with potassium amide. Observations are as follows:

(*a*) When reaction is completed, the evolution of gases ceases and the solution has a pale yellow color. Presently, the color darkens and, in the course of 10 to 20 minutes, carbon precipitates. Approximately 60 per cent of the total carbon appears as free carbon. The residue in the tube consists of carbon and potassium chloride and the weight of the contents of the reaction tube is less than that of the reactants.

(*b*) The gas collected during reaction is nitrogen. Somewhat less than one atom of nitrogen is obtained per atom of precipitated carbon.

(*c*) A volatile carbon compound is produced which decomposes as it is carried into water in collecting the evolved gases. Flashes of light appear in the collecting tube as the bubbles of gas, largely ammonia, collapse. A scum of carbon forms on the surface of the water in the collecting tube.

(*d*) In one experiment, the ammonia was taken off at low temperatures; when the tube was allowed to warm up there was a mild detonation accompanied by a flash of light and carbon was deposited on the interior walls of the reaction tube. In this case, there were no flashes in the absorption tube and no carbon appeared on the surface of the water in the collecting tube.

The reaction of hexachloroethane with potassium amide is deep-seated and complex. It seems likely that the evolved nitrogen is due to the decomposition of some compound which leaves free carbon, on the one hand, and produces nitrogen, on the other. Roughly 35 to 40 per cent of the total carbon forms an unstable compound which is fairly volatile. It decomposes in the collecting tube to form carbon and, perhaps, nitrogen. No other gas appears. This compound appears to be unstable at room temperature and decomposes explosively.

The above reactions serve to illustrate how little we know about the chemistry of liquid ammonia and how greatly it differs from that of other solvents.

It would be of interest to know what would be the reaction of potassium amide with highly halogenated derivatives of aliphatics containing more than two carbon atoms per chain. It would also be of interest to know what reactions might occur in the case of highly halogenated aromatics, particularly hexachlorobenzene.

Solutions of the alkali metals in liquid ammonia are much more powerful reducing agents than is potassium amide. However, in order to be able to interpret such reduction reactions, we need to have some knowledge of the reaction of the amides. In the reaction of the metals with alkyl halides, the metal amide is formed which, in turn, reacts with the halide to form amine. Thus, sodium reacts with methyliodide to give a quantitative yield of methane and methylamine in equimolar proportions.

We know nothing about the reactions of alkali metals with polyhalides. A study of such reactions might throw much light on reactions such as those outlined above.

DR. WILL S. DELOACH, author of "Doctorates in Chemistry from Southern Institutions, 1881–1951" on page 190 of our April issue, has recently learned of two additional chemistry doctorates which should have been included in his list. They were granted by the Agricultural and Mechanical College of Texas in 1939 and 1951.

Proceedings of the

PACIFIC SOUTHWEST ASSOCIATION OF CHEMISTRY TEACHERS

SALESMANSHIP IN CHEMISTRY TEACHING[1]

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You can't sell chemistry as a teacher unless you are sold on teaching!

You and I have a commodity to sell. Do we make the most of our opportunity and obligation? It is my conviction that a chemistry teacher must use salesmanship just as skillfully as do the book-men who call upon us, or the representatives of the apparatus companies. They make us want what they've got even if we have to go through the painful operation of writing and justifying requisitions to get it.

We chemistry teachers have a tremendous advantage and challenge. The advantage exists because our subject has found its way as a requirement in our college curricula. Our subject has been popularized as the science that has provided "better things for better living." Consequently, when we signed on at our respective schools to teach chemistry we knew that there would be classes waiting for us on the opening day of school. We also knew that, in spite of us, many of our students would stay on until the end of the term. They had to have credit in chemistry and they were determined to get that credit regardless of the mixture in the dish that we set before them.

All too frequently the courses set before our young men and women are like inexpensive pre-packaged ice cream—full of froth and flavor, without enough body for the aspiring young scientists to set their teeth into. It is far better that they be given a heavy hash of problems, equations, and theory; those who survive such a course will at least have nourishment. But it seems to me that to be truly effective our courses and textbooks should provide a balanced menu. There should be a scientific *hors d'oeuvre* to whet the appetite; the *entré* must be of solid demonstratable facts; the salad, of vitamin-rich theory to insure proper growth; well-prepared garden-fresh problems and questions will serve as vegetables for this meal; for the student who has time and capacity it is well to provide a pie-a-la-mode of outside reading and special assignments and experiments. Yes, I also consider it very wise to supply an occasional sip of humor, for the wine of laughter is a great stimulus to mutual respect and understanding. It is well that our students be thoroughly satiated at the close of the hour in our classrooms so that they will have plenty to digest before they come again.

When we provide this sort of menu our students will be sold; they will be glad to come again to another feast. At the end of the term when they leave our classrooms they will anticipate the more rigorous fare of advanced courses and individual creative work.

From the point of view of the junior college and the liberal arts college the problem is considered to be different from that of the university with a large graduate department. It would be wise to grant the existence of a difference in degree only. Wherever chemistry is taught it should be done with the greatest skill and salesmanship. The realization of this fact would result in more effective teaching in all types of institutions. A greater number of promising students would select science, and specifically chemistry, as their major, thus insuring our nation of an ever available supply of technically trained personnel.

In order to reach some concrete conclusions regarding the selling of a serious interest in chemistry to our students, let us evaluate four areas of influence: (1) the college, (2) the pedagogic technique, (3) the subject matter, and (4) the teacher.

Since ideas regarding the relative merits of these four items are obviously somewhat controversial I will present my views in the hope that they will at least provoke thought, to the end that each one of us will strive to teach better. (The preparation of this short paper has already brought into focus a number of respects in which I have not been doing as effective a job of selling chemistry as I should.)

Let us then consider the institutions that do the best job of selling science to their students.

In the July, 1951, *Scientific American* is an article entitled, "The Origins of U. S. Scientists." Heading the list of 50 institutions whose male recipients of bachelors' degrees have gone on to obtain the Ph.D. in natural science, is Reed College in Oregon with the amazing number of 131.8 out of each thousand graduates. In vain I scanned the list for Northern California's two great universities. The fact that they were not

[1] Presented before the PSACT on November 3, 1951.

included was softened however by the discovery that our California Institute of Technology at Pasadena was second on the list with 70.1 per thousand graduates. Note that this is only slightly more than half of the number for Reed. Then I sought my Alma Mater, finding to my surprise that two small colleges, one to her north and one to her south, ranked 25th and 32nd respectively, but Utah was absent. I further noted with satisfaction that the pioneer spirit of the West plays its part, for a far higher percentage of the possible schools of the West made the list in comparison with eastern colleges and universities.

Space will not permit extensive discussion of this article. It is significant, however, to quote briefly regarding two points. In discussing Reed College the article states:

> Since its founding in 1911 Reed has had a brilliant record of achievement. During 1925 to 1940 it has produced 12 Rhodes scholars... Though salaries (at Reed) have been relatively low, many top-notch men have come to Reed to teach and have stayed there, disdaining more lucrative positions. Among the students, the campus hero is the scholar. The curriculum is organized to foster a maximum of individuality of instruction, and teachers and pupils alike carry on a tradition of disputatiousness which in many institutions might be a sign of disorganization and dissatisfaction.

The article furthermore describes the type of instructor who is most effective:

> Statistical examination of the ratings assigned by students and by the investigator indicated first that a successful teacher is not especially distinguished for his mastery of superficial pedagogic skills. Rather, the successful teachers are marked by three cardinal traits: masterfulness, warmth and professional dignity. It would appear that the success of such teachers rests mainly upon their capacity to assume a father role to their students, in the best sense, and to inspire them to an emulation of the teacher's achievements.

The conclusions reached in this article indicate that the most fertile soil in which to set the feet of a young man who would be inspired to scientific achievement is found in the smaller liberal arts colleges. One obvious question is "Why?" May I be so bold as to suggest the answer as "inspiration rather than mass-production"?

So much for the colleges! Second, let us glance at the pedagogic techniques to be employed in the selling of chemistry to our students. Extensive training in methods of teaching is not essential. In fact, it is notable that the presidents and most of the faculty members of most of our great universities and liberal arts colleges have never had such courses. According to school law these men are not qualified to teach in our secondary schools, never having plowed through the mumbo-jumbo of the required courses in education. On the junior college and high-school level there seems to be the conviction that nearly any credentialed individual is qualified, through the divine guidance of courses on the techniques of teaching, to present the subject matter of any course. Yet every teacher knows that there are successful teachers who have never been exposed to courses on how to teach. All of us have developed many of our own techniques through experience as the need arose, and through the emulation of outstanding men under whom we have studied.

In chemistry instruction there are certain classroom techniques that have become traditional. They are typical sales techniques. The salesman spreads his wares before the prospective buyer so that the buyer will see, admire, and desire. We have all been influenced by the spread put out on our living room floor by the Fuller Brush man. We know that our students are influenced by the effective use of demonstrations, charts, chemical displays, films, and apparatus set-ups. Chemistry teaching, to be most inspirational, needs these visual aids. They should be relatively simple, easy to follow and to understand. They may be dramatic. Fundamentally, however, they must illustrate scientific principles and clarify facts.

Do you and I use these visual devices to their best advantage? From a consideration of the department in which I serve, I know that we do not. I suspect that most of you do not. It takes a lot of time and effort to prepare effective demonstrations for a chemistry class. In this respect the professors of our large universities have a tremendous advantage because they are provided with the services of student assistants; many teachers in smaller colleges are similarly fortunate. Because of heavy teaching loads and lack of adequate facilities, we in the junior college and in most state colleges fall far short of making the most effective use of these devices. Yes, and some of us are too lazy. Most of us could do better. It is my conviction that every chemistry lecture should include some demonstration. Can we convince our boards of education that we need time allowed in our class schedules for preparation?

An opportunity for closer association with our students is the advantage of the smaller schools. A chemist acquaintance of mine who is well on his way to his first million states, "I wasn't a very good student, but one of my profs gave me all the personal aid that I needed. He is largely responsible for my success." Contrast that experience with the lecture section of 500 broken into 25 lab groups which are supervised by student assistants whose own superior ability in chemistry makes them unsympathetic with the difficulties of their charges.

Discussion on the points of pedagogic techniques could go on indefinitely, but such discussion is not necessary. Each of us can afford to do a little self-evaluating on this point. Perhaps we can diplomatically influence our colleagues!

Third, let us focus our attention on subject matter. In whatever chemistry course we teach, the essential subject matter of that course should be covered. It was with great regret that I was unable to attend the meetings of our organization when the topic for discussion was "What can we throw out of the course in general chemistry?" Knowledge about chemistry is so voluminous that it is easy to lose sight of the fundamental principles of the science in the unlimited maze of facts through which we must strive to lead our students. Our textbooks have become so large that

they are cumbersome and often discouraging. I have often felt that they have been written with the motive of showing how much the author knows rather than with the desire to impart fundamental knowledge to the student.

It is my conviction that we in chemistry must critically select the fundamental principles that are essential to the courses that we offer, and that we must use these principles as the basis of our teaching. We cannot hope to cover all of the related facts. Facts should be drawn from as wide as possible a range of applicable industrial, social, and historical data, but their function should be to establish the few fundamental principles. If we can successfully weed out the extraneous subject matter we will be more successful as teachers. It is notable that western institutions have taken the lead in teaching "principles."

It is poor sales psychology to expect the student to carry around a textbook that is obviously too large to be mastered during the semester or the year allotted for the course. One great disadvantage of larger books is that the teacher must skip-assign; this cannot be done without the loss of continuity in subject matter and in the scientific philosophy that is being developed.

It is also poor sales psychology to require the student to buy a textbook in which assignments are not made. One freshman chemistry teacher, a user of the splendid Frantz lab manual, said to me, "I find this manual entirely adequate for the course. I don't make three assignments in our selected lecture text during the whole term." Then I heard a student comment, "Yea! Then why'd he make us buy the darned thing—it cost $4.75."

Another disadvantage of the large text is that the conscientious teacher and his students become bound to it in the attempt to justify its purchase. This removes the teacher's greatest opportunity for salesmanship—the privilege of inspiring the student by drawing from personal experience and from the splendid storehouse of chemical literature to enrich his course with current factual information that illustrates the principles that he is teaching.

The literature of the field of chemistry, *i. e.*, articles to be found regularly in the various periodicals are excellent reservoirs of inspiration to use in selling interest to the general student.

Finally, let us consider ourselves, the teachers who do the selling. My opening sentence was, "You cannot sell chemistry as a teacher unless you are sold on teaching." Stated otherwise, one can best sell an interest in chemistry if he is sincerely interested in teaching, if teaching chemistry is his first and greatest interest.

The effectiveness of our teaching profession is threatened by a condition that has become alarmingly prevalent in the last few years. Many potential M.D. and D.D.S. candidates, unable to secure admittance to medical and dental schools, are deciding to become high-school and junior college science teachers, saying, "No, I don't particularly care for teaching but . . . ." There are numbers of such young people attending our State Teachers Colleges. They are capable young people but for them teaching is a compromise. Let us hope and pray that someone with whom they come in contact will be able to make teaching become truly their first love, and not just a stop-gap job. If they are to be sold on the vocation of science teaching as one of the most worth-while tasks it must be done by science teachers. I cannot conceive of that inspiration being gained from courses in education! We who are their science teachers must show them by example that our vocation is worth while and that we love our work. We must always play the game.

The standing of our profession is frequently weakened by our colleagues who do not play the game. I recall the students' lack of respect for a teacher who would not admit and correct an error in his explanation; the disgust for a biology teacher who doped every solution to be sure of getting a digestion reaction; the fear of a teacher who belittled the capacity of his students; the hatred of a teacher who consistently drove the poorer students from his classes and embarrassed those remaining; the pity for a teacher who constantly fumbled because he did not know his subject matter; the indifference to the teacher who was bored with his job. One final criterion of the success of a chemistry teacher is that he has not only inspired the chemistry majors, but that also the general students in his classes have found chemistry to be an approachable and understandable subject, worthy of the student's respect.

The salesman who comes to your door must possess certain personal traits. He must be diplomatic. If he expects to come again for a repeat sale he must be honest. He must display versatility, adaptability, and resourcefulness. To be really successful he must know his commodities thoroughly and must believe in their value. Above all, a salesman must never lose sight of his objective, the signature on the dotted line.

So it is with us. A chemistry teacher must never lose sight of his objective—the absorption by the student of knowledge and of inspiration to such an extent that he, the student, may be better equipped to face the future. If we as teachers do not give our students both knowledge and inspiration we will be wasting much of our time and theirs. When we do give both knowledge and inspiration the rewards to us, although often intangible, are rich indeed.

Report of the

NEW ENGLAND ASSOCIATION OF CHEMISTRY TEACHERS

OFFICIAL BUSINESS

MEETINGS

The 264th Meeting, held at Brown University, Providence, Rhode Island, February 9, 1952, was opened by Leslie B. Coombs, Chairman of the Southern Division of the NEACT, who introduced Dr. Samuel T. Arnold, Provost of the University. In welcoming approximately 150 members and guests to this meeting, Dr. Arnold evoked concern for the prospective shortage of science students for the near future, and suggested that teachers of science, especially in the secondary schools, had a grave obligation of guiding good students into scientific fields.

Dr. James S. Coles, Acting Dean of the College, brought a progress report on the new Bachelor of Science program in chemistry at Brown University. He outlined briefly the program of the first two years at Brown which includes a year of the chemistry of the covalent bond, a year of physical chemistry, and a semester of inorganic chemistry in the second half of the sophomore year. During the junior year a year of analytical chemistry, having physical chemistry as a prerequisite, is required of every candidate. Then in the senior year students are required to take one course each from the fields of organic chemistry, inorganic chemistry, and physical chemistry, and may elect other courses in these fields as well as in biochemistry and industrial chemistry. He suggested that one of the big advantages in this program is the large initial enthusiasm which students from high school bring to chemistry which is ordinarily not manifested if the first course in chemistry is just the "same old stuff" that they had in high school. In assaying the program to date, Dr. Coles and the staff at Brown feel that students in this curriculum have done as well in all courses as they did in the previous program at Brown, in spite of the fact that qualitative analysis is no longer given. Another feature of the program is that physical chemistry is not given piecemeal in courses in general chemistry, and qualitative analysis and quantitative analysis, and then repeated again in a course in physical chemistry as it was in actuality under the old program.

Mr. Robert M. Sherman, instructor at Brown University, then gave some demonstrations for beginning chemistry in which various principles were illustrated. His demonstrations were characterized by simplicity of apparatus, and he emphasized that having the student see a single thing in a demonstration is the best possible means of presentation. Copies of directions for carrying out these demonstrations were distributed.

After a luncheon at the new Sharpe Refectory at Brown University, a short business meeting was held at 1:30 P.M. Dr. Harold C. Harrison, Associate Professor of Chemistry, University of Rhode Island, spoke on the subject "Scientific crime detection." Dr. Harrison gave a short history of the introduction of scientific methods into crime detection, and then demonstrated some of the ways, often simple, in which chemistry may be used for revealing after the fact events which have taken place at a crime.

The Secretary reported that membership had reached 516 members, which includes eight new members since the December meeting. Five resignations have been received and 24 have been suspended for nonpayment of dues. The new members welcomed into the NEACT at the meeting were:

LeRoy Bachman, 24 King Street, Pottstown, Pennsylvania.
Louis C. W. Baker, Chemistry Department, Boston University, Boston, Massachusetts.
John S. Belew, Department of Chemistry, Brown University, Providence, Rhode Island.
Robert S. Bigelow, Teachers' College, Farmington, Maine.
Alexander M. Cruikshank, 1232 Kingstown Road, Kingston, Rhode Island.
William T. Macdougald, 19 Catalpa Road, Warwick, Rhode Island.
Melvin Merken, Old Lyme School, Old Lyme, Connecticut.
S. Ricklin, Department of Chemistry, Brown University, Providence, Rhode Island.

The 265th Meeting was held at Deerfield Academy, Deerfield, Massachusetts, April 12, 1952. Dr. Frank L. Boyden, Headmaster of the Academy, opened the meeting with a short historical sketch of some of the people of Deerfield early connected with the founding of the Academy. Professor Anna Jane Harrison described the freshman chemistry course at Mt. Holyoke College entitled "Basic Chemistry." It is a laboratory-centered course, almost devoid of descriptive chemistry not requiring high-school chemistry as a prerequisite. The second paper in the morning, "Interesting products and new developments of the General Electric Chemicals Division," was given by Dr. Kurt C. Frisch. He described plastics, silicone resins, and other new products which have recently been developed

by the Chemicals Division, and illustrated many of them. A scrumptious dinner was served by the boys of the Academy, climaxed by a frozen dessert in a flower pot, the latest thing according to *House and Graden* magazine. A short business meeting followed. Professor C. Pauline Burt of Smith College concluded the meeting with a talk entitled, "A survey of synthetic medicinals." The rate of progress in this field is demonstrated by the fact that Dr. Burt gave a similar talk under the same title before the 1944 Summer Conference and in her 1952 talk did not touch any group of medicinals mentioned in the earlier one, except the sulfa drugs.

Many of the group of 100 who gathered for this meeting visited some of the 21 prerevolutionary, historic houses of Old Deerfield after the formal program of the day.

The Secretary welcomed four new members to the NEACT, making the total 526:

W. Gordon Brown, Principal Senior School, Hopkins Grammar School, New Haven, Connecticut.
Louis J. Bush, Head of Chemistry Department, Greenfield High School, Greenfield, Massachusetts.
Curtis P. Hinckley, Woodstock Country School, Woodstock, Vermont.
Sister Ernestine Marie, Ryan Memorial High School, Dorchester, Massachusetts.

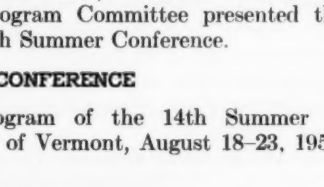
William F. Ehret, Chairman of the Summer Conference Program Committee presented the program for the 14th Summer Conference.

SUMMER CONFERENCE

The program of the 14th Summer Conference, University of Vermont, August 18–23, 1952, is as follows:

Aug. 18, 1952, Evening Session
Address of Welcome by Professor Charles E. Braun, Chairman of the Department of Chemistry, University of Vermont.
"Recent research on sugar maples," Professor James W. Marvin, Chairman of the Department of Botany, University of Vermont.
Aug. 19, 1952, Morning Session
"The newer synthetic fibers," Dr. Joseph B. Quig, Manager of the Development Section, E. I. du Pont de Nemours and Co.
"The silicones and their neighbors," Professor Eugene G. Rochow, Department of Chemistry, Harvard University.
Aug. 19, 1952, Afternoon Session
"Psychological factors involved in taking Science Talent Search examinations," Dr. Harold S. Edgerton of Richardson, Bellows, Henry and Co., Inc.
"Titanium—metallurgy and industrial applications," Dr. C. Kerby Stoddard, Development and Engineering Department, National Lead Company, Titanium Division.
"Some recent developments in steel manufacture and applications," Mr. William A. Dennis, United States Steel Company.
Aug. 19, 1952, Evening Session
"Criminalistics," Professor Mary L. Willard, Department of Chemistry, Pennsylvania State College.
Aug. 20, 1952, Morning Session
"The behavior of metal complexes in aqueous solutions," Professor Arthur E. Martell, Department of Chemistry, Clark University.
"Chemical reactions in corrosion of metals," Dr. Robert M. Burns, Chemical Coordinator, Bell Telephone Laboratories.
Aug. 20, 1952, Afternoon and Evening
Chicken barbecue and whist party
Aug. 21, 1952, Morning and Afternoon Sessions
"The use of available radioisotopes in the teaching of general chemistry"; demonstrations and lectures to be given by representatives of the United States Atomic Energy Commission.
Aug. 21, 1952, Evening Session
"Science in Asia," Dr. Robert R. Williams of the Research Corporation.
Aug. 22, 1952, Morning Session
Symposium: The Evaluation of Laboratory Work in General Chemistry
Professor Arnold J. Currier, Department of Chemistry, Pennsylvania State College.
Mr. Carl Swinnerton, Pomfret School, Pomfret, Connecticut.
Professor Andrew J. Scarlett, Department of Chemistry, Dartmouth College.
Aug. 22, 1952, Afternoon Session
"Chemistry in the food industry," Dr. John H. Nair, Assistant Director of Research, Thomas J. Lipton, Inc.
"Whys and wherefores of vocational guidance in chemistry," Dr. Walter J. Murphy, Editor of *Chemical and Engineering News*, *Industrial and Engineering Chemistry*, and *Analytical Chemistry*.
Aug. 22, 1952, Evening Session
"The first fifty years are the hardest," Mr. Augustus Klock, The Fieldston School, Bronx, New York.

For information, write to the Conference Secretary, Miss Constance L. Brown, Department of Chemistry, University of Vermont, Burlington, Vermont.



To the Editor:

May I add the following remarks to the recent paper by Dr. C. D. Coryell "The periodic table: the *6d-5f* mixed transition group" (J. CHEM. EDUC., **29**, 62 (1952)).

The so-called actinide hypothesis was first proposed by Dr. G. E. Villar of the Facultad de Ingenieria de Montevideo, Uruguay, as early as 1938 (*Bol. Fac. Ing. Montevideo*, **5**, 231 (1938); *Chem. Abstracts* **33**, 2003–4 (1939)). Later he published detailed chemical, physical, and crystallographic data supporting his actinide hypothesis (*Ann. Acad. Brasil. Sci.*, **12**, 51 (1940); *An. Ass. Quim. Argentina*, **31**, 213 (1943)). A short account of this hypothesis was also published in the United States (J. CHEM. EDUC., **19**, 286 (1942)).

Villar's actinide hypothesis was confirmed by the brilliant researches of Seaborg and associates at the University of California.

I should like to point out that Glockler and Popov (J. CHEM. EDUC., **28**, 212 (1951)) were not the first to propose that the neutron ($_1{}^0n$) may be included as the element of atomic number 0, as stated by Dr. Coryell. Although I have failed thus far to find the originator of the idea, many Brazilian and Spanish chemical textbooks include this hypothesis. In 1949 Dr. E. Silva proposed not only to consider the neutron as the element of atomic number 0 but also to consider it as the first noble gas (SILVA, E., "Quimica inorganica," Recife, **1949**).

RICARDO CARVALHO FERREIRA

FACULDADE DE CIENCIAS MEDICAS
RECIFE, PERNAMBUCO, BRAZIL

• PROBLEM BOOK FOR GENERAL CHEMISTRY

Royce H. LeRoy, **Professor of Chemistry, A. & M. College of Texas. McGraw-Hill Book Co., Inc., New York, 1951. v + 170 pp. 7 tables. 21.5 × 28 cm. $1.90.**

THERE are 26 sections to this book of problems in general chemistry. The commonly accepted categories such as "Percentage composition from the formula" and "Equivalent weights and normal solutions" are preceded by "Arithmetic review" and "The metric system of measurement." Most of the sections contain from 10 to 26 problems, but the last is a general review of 57, and a rough check indicates a total of 427 problems in the book.

Use of letter-size pages, $8^1/_2 \times 11$ inches, with only five or six problems to the page, leaves plenty of room to write solutions in the book itself. The completed pages can be easily removed and reassembled by means of holes already punched.

One of the easiest of the questions is: "What is the volume of 500 g. of concentrated sulfuric acid?" One of the hardest questions is: "One drop of 0.1 *M* hydrochloric acid was added to 20 ml. of a saturated silver chloride solution. Ignoring the slight increase in volume, calculate (*a*) the total resulting chloride concentration and (*b*) the concentration of the silver ion remaining in solution in equilibrium. (*c*) What weight of AgCl should have precipitated?" Incidentally, the answer given in the back of the book to part (*a*) makes use of an approximation that gives an answer to part (*c*) that is quite inaccurate.

Each section is preceded by a page or so of explanation. Most of the explanatory text is scientifically accurate, but the explanation of equivalent weight, for example, is in need of revision.

There are at least two educational reasons to justify the use of a book of this kind. For the first, it should help the busy instructor to vary assignments so that students of different capabilities can be trained in the techniques of solving particular kinds of chemistry problems. Although this book provides plenty of challenge for some, and additional practice for others, there is little to hold the interest and attention of a first-rate mind.

Many believe, moreover, that in addition to providing training in the techniques of solving particular problems, a book of this kind should provide opportunities to learn to apply the generalizations of chemistry in solving difficulties presented by situations that are almost entirely new. In the book under review there is no hint of the recognition of this fundamental aim of all science education. Instead, there seems to be only a recognition of the need to develop stereotyped and, therefore, superficial thinking. Strengthening the impression that the solution of these particular problems is to be but an end in itself, we find that all generalizations are clearly handed to the student. He is never permitted to wade through a bog of data in order to enjoy the thrill of capturing for himself even the most elementary principle. The author states the more restricted point of view most clearly in the final sentence of his preface, when he says, "It is the hope of the author that he has succeeded in evolving an orderly approach to the problems of chemistry based on the simple principles of arithmetic, and that the student will work his way successfully through the book from beginning to end."

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• TOPICS IN PHYSICAL CHEMISTRY

W. Mansfield Clark, **DeLamar Professor of Physiological Chemistry, The Johns Hopkins University School of Medicine. Second edition. The Williams & Wilkins Co., Baltimore, 1952. xviii + 777 pp. Illustrated. 16 × 23.5 cm. $10.**

MANY, if not most, of the physical chemistry textbooks designed for the use of premedical and medical students are simply diluted versions of the material studied by students majoring in chemistry. Also, these texts often apply the principles of physical chemistry to the solution of biological problems as an apparent afterthought. As a result, the students believe that the science of physical chemistry is too impractical to be applied to the study of the art of medicine. This book does not suffer from these deficiencies. Admittedly, it does not fully cover all of the classical subdivisions of the subject but this is not necessary in order to accomplish the purpose intended. For example, treatments of the solid state and of the phase rule are not to be found in this text.

Briefly, the author selects topics of interest to the student and illuminates them in the strong light of physical chemistry. From experiences with the first edition, the student will come to look upon physical chemistry as a powerful friend instead of an indifferent and obscure entity. The chapter headings, abridged, are: Limiting laws, review of conventions, the balance, measurements of volume, density, sedimentation, the gas laws, colligative properties, distribution between phases, diffusion, semipermeable membranes, kinetics, equilibria, electrolytes, conductance, pH, protein solutions, equilibria in the blood electrolytes, thermodynamics, redox relations, potentiometry, the polarograph, atomic and molecular structure, isotopes, refraction and polarization, stereoisomerism, radiant energy, luminescence, colloids.

This book, written as a supplementary text, is too voluminous to be covered in a one-or even two-semester course. Nevertheless, it can be used as a text for premedical physical chemistry

by selection of subject matter and with the use of other reference works. In addition it is written with sufficient clarity, in an almost conversational style, so that it can be used later by the student with little or no outside help, once he has been introduced to the subject. It is therefore a good reference work for the medical student and for the graduate biological student.

The author approaches the use of calculus in a practical manner. As he states at length in the appendix, calculus is necessary; it is after all just common sense expressed in unfamiliar symbols—not really difficult—so let's approach it with confidence. In this connection, the book does not have enough problems to teach the student how to apply mathematics to real problems. This is probably the outstanding weak point of the book. Perhaps future editions will be different.

For the inquisitive student many references to review articles and to the original literature are cited immediately after the discussion of the point at hand or in the body of the discussion. General references are listed at the end of the chapters.

Changes in this edition are, for the most part, either improvements in the already clear explanations or are revisions to bring the discussion up to date. In particular, the chapter on atomic and molecular structure has been almost completely rewritten. Other changes of lesser magnitude have been made, for example, in the chapters on electrolytes, pH, protein solutions, thermodynamics, redox relations, and colloids. An additional change would have been to include a section on photochemistry; it is unfortunate that no such change was made.

The printing, binding, and typography are satisfactory. No significant typographical errors were noted. There were no observed errors of fact or interpretation, provided it is remembered that this is an introductory text for students of medicine.

This book does not make physical chemistry easy for the premedical or medical student. It does make it evident that an understanding of physical chemistry is necessary for a good grasp of the subject of biology, and secondly it makes this study interesting. More cannot be asked of any text.

JAY A. YOUNG

King's College
Wilkes-Barre, Pennsylvania

● MODERN TRENDS IN PHYSIOLOGY AND BIOCHEMISTRY

Edited by *E. S. Guzman Barron,* Chemical Division, Department of Medicine, University of Chicago. Academic Press, Inc., New York, 1952. xxii + 538 pp. 132 figs. 15 × 24 cm. $8.50.

The 1950 physiology course of the Woods Hole Marine Laboratory was intended to be dedicated to Professor Leonor Michaelis in commemoration of his 75th birthday. Due to his death in September, 1949, this plan could not materialize, but the lectures presented at the course have been collected in this memorial volume and they represent an eloquent tribute to a great scientist, whose fundamental contributions are stressed in nearly every one of the 20 chapters. The book contains a good portrait of Michaelis and the introductory chapter is the memorial lecture read by Barron at Woods Hole.

The book is difficult to review briefly. Its scope ranges from pure biochemistry and biophysics to the morphological basis of function. The presentation varies from a vividly written, essay-like chapter with four references to strictly technically presented chapters with up to 177 references. The chapters range in length from 11 to 42 pages, and some contain no figures while 40 figures are found in one.

The editor points out that the topics covered can be divided into six groups: cellular physiology (five chapters by Runnström, Mazia, Heilbrunn, Parpart and Ballentine, and Jacobs, dealing, respectively, with physiological aspects of cytoplasm, cell nucleus, and cell division, with molecular anatomy of the red cell membrane and cell permeability); muscle physiology (three chapters by Szent-György, Mommaerts, Kupfler, who discuss thermodynamic questions, the contractile protein systems, and transmission processes at nerve-muscle junctions); nerve physiology (five chapters by Steinbach, Grundfest, Nachmansohn, Tobias, and Kupfler's chapter already mentioned under muscle physiology. Discussed are the sodium and potassium balance of nerve and muscle, bioelectrical potentials, chemical mechanisms of nerve activity, ultrastructure and function); enzyme action (five chapters by Barron, Chance, Klotz, Neurath, the former discussing in 2 chapters the mechanism of enzymatic oxidation-reductions and the oxidative pathways of carbohydrate metabolism, while the other authors present one chapter each on enzyme-substrate compounds, metal complexes with proteins, and chymotrypsin); mode of drug action (one chapter by Loewi); and finally biochemical evolution (one chapter by Wald).

The list of authors, each a recognized authority in his field, is a sufficient guarantee that everybody interested in any of the above topics will find their contributions well worth reading. It should nevertheless be pointed out that the title of the book is somewhat too broad. "Some modern trends in physiology and biochemistry" would have characterized the book better than the chosen title. As already indicated, bibliographies of varying length are attached to each chapter. The book contains an author and a subject index and is very well printed.

THEODOR VØN BRAND

National Institutes of Health
Bethesda, Maryland

● COPOLYMERIZATION

Turner Alfrey, Jr., Dow Chemical Co., _John J. Bohrer,_ International Resistance Co., and _H. Mark,_ Polytechnic Institute of Brooklyn. Volume VIII of "High Polymers." Interscience Publishers, Inc., New York, 1952. x + 269 pp. 77 figs. 41 tables. 15.5 × 23.5 cm. $6.80.

Copolymerization is the production of high molecular weight compounds in which two or more different monomer units alternate in the chain molecule. Such materials frequently have useful properties superior to those of polymers made from a single monomer. The technological importance of copolymers (synthetic rubbers, plastics, and fibers) has led during the last decade to a large amount of both experimental and theoretical work on the principles of copolymerization and its timely summary will be of great assistance to those working in this field.

The book contains detailed discussions of the kinetics of copolymerization by the free radical mechanism, the factors influencing monomer reactivity, the behavior of divinyl monomers and the vulcanization of polymers obtained from them, copolymerization by ionic mechanism and the physical properties of copolymers. The treatment is much more complete than that of Mayo and Walling's survey (*Chemical Reviews,* **1950**) and it should be easier reading for those who have had little previous experience in this field. The authors are very scrupulous in presenting all views on controversial subjects in a fair and objective manner. The extensive discussion of the estimation of errors in the calculation of monomer reactivity ratios is a useful reminder of the limitations imposed on the interpretation of experimental copolymerization data. A table listing monomer reactivity ratios for 275 monomer pairs is an invaluable reference.

Different chapters were apparently written by different authors and this resulted in some repetition. The absence of any effect of the dielectric constant of the solvent medium on monomer reactivity ratios is accepted on page 68 as a valid argument against the derivation of the Alfrey-Price theory, while the argument is refuted on page 114.

The book is a valuable addition to the well-known "High Polymers" series and is to be recommended whole-heartedly to those engaged in industrial research and development of high polymers and those interested in academic studies of polymerization kinetics.

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By I. M. KOLTHOFF, *Professor and Head of Division of Analytical Chemistry, University of Minnesota, and* JAMES J. LINGANE, *Professor of Chemistry, Harvard University.*

The publication of the first edition of this book inaugurated a new era in polarography. Numerous new fields of application were introduced, and a deeper understanding of the theoretical fundamentals led to a more critical approach to experimental results. The literary output of this "polarographic movement" resulted in tripling of the published literature. It needed the superior guidance of I. M. Kolthoff and James J. Lingane, aided by Stanley Wawzonek in the field of organic polarography, to present again the wealth of material in a well-organized and integrated unit. *Polarography* will thus remain the definitive treatise for every one who must use this type of measurement in scientific and industrial investigations.

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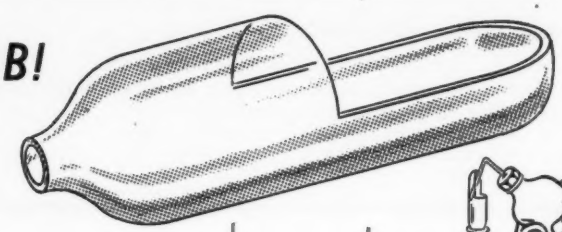

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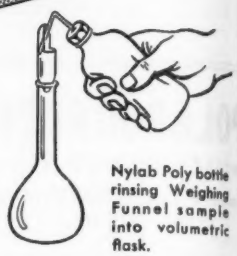

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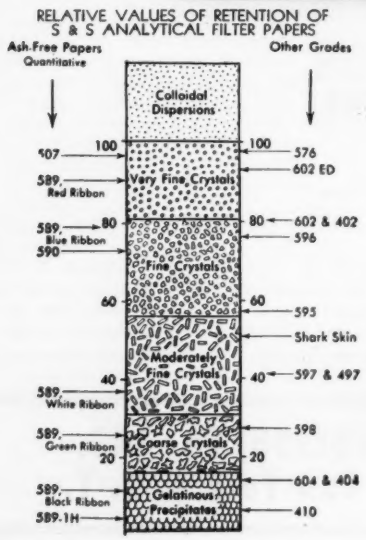

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Differential Colorimeter

Manufacturers Engineering & Equipment Corporation, Hatboro, Pennsylvania, announces the "Colormaster Differential Colorimeter," a useful instrument for obtaining accurate color measurements of synthetic fibers, paints, liquid solutions, pigments, slurries—any substance, in fact, where color is an important factor.

According to the manufacturer a complete set of color measurements can be made with the Colormaster in less than thirty seconds. The new instrument has a system sensitivity appreciably greater than the eye can discern, and is particularly useful for measuring differences between similar colors for process control applications. It is extremely simple to operate, and results are taken directly from an easy-to-read dial.

Three color filters give approximations to the I.C.I. X, Y, and Z functions, and transmissions and reflections are measured directly. The linearity of the instrument is within approximately 0.2 per cent reflectance or transmission down to 10 per cent; by standardizing on a known dark panel, the range can be extended far below 10 per cent, with linearity within approximately 0.02 per cent. The dial can be read to 0.1 per cent, and hundredths of a per cent can be readily estimated. The dead zone (back-lash) does not exceed 0.02 per cent reflectance or transmission. Because the Colormaster is highly stable after a warm-up of about one-half hour, it requires infrequent standardization. Magnetic shielding is included to permit measurement of colors where steel panels may be used. The electrometer input impedance is about 6000 megohms. Two calibrated, highly stable white panels are furnished with the instrument.

Beckman Custom Instruments

Beckman Instruments, Inc., has recently established a new Special Products Division which will be devoted primarily to the study and development of special instruments brought to it by industrial organizations and educational institutions. Need for such a division has become increasingly apparent in recent years with the rapid advancements in instrumentation. Many such problems have been brought to the Beckman organization for study and possible development, but due to the large-scale nature of the normal Beckman Instrument operations, it has been difficult to give adequate consideration to many of these instrumentation problems — particularly those requiring special production facilities or development programs.

The newly-established Beckman Special Products Division is a completely separate organization within the Beckman Instru-

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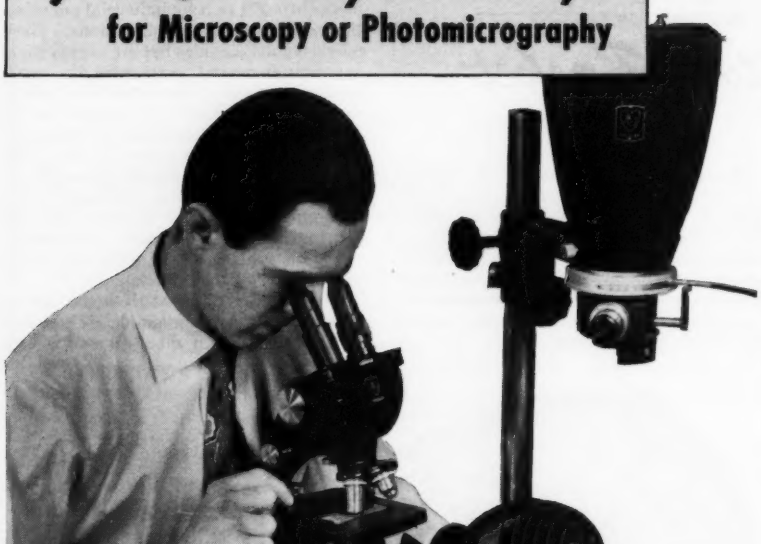

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laboratory news

A new monochromator is now offered which eliminates the elaborate paraphernalia of arc light and prism spectroscope to secure monochromatic light. The new monochromator employs an ordinary tungsten light source and uses the rotary power of quartz. It will be welcomed for microscopy and other purposes which require monochromatic light.

Not only does the monochromator represent simplified equipment but it gives an increased intensity of light as well. Collimated light from a tungsten filament lamp will yield adequate illumination at all wave lengths because of the large aperture available — a full inch.

In this quartz monochromator, white light from the tungsten source is passed through basal sections of optical quartz laminated to polaroid sheets. By a synchronous rotation of the quartz sections, the wave lengths of the transmitted light can be varied continuously through the visible spectrum. The instrument produces an average band width of 150 Angstrom units. A built-in adjustable support is provided for tilting the device in line with the light source and the microscope or other apparatus to be used.

The prism spectroscope uses a space-taking, cumbersome, carbon arc light source. It requires direct current and a mechanism to keep carbon points in adjustment. Light passing through prism yields as little as .5% beam on slide.

The new quartz monochromator uses ordinary inexpensive tungsten filament lamp as light source. No maintenance is necessary — bulbs are changed quickly and easily.

For prices and delivery of the monochromator, write Cambridge Thermionic Corporation, 476 Concord Ave., Cambridge 38, Mass.

ment operations, and is located in its own modern plant, with complete development and production facilities particularly designed for the manufacturing and testing of custom-built instruments. At the same time, the full experience and "know-how" of the Beckman organization in advanced electronics and modern instrument design will be at the disposal of the new division, thus combining the valuable Beckman experience and facilities in large-scale instrument production with the versatility and individual attention characteristic of a small centralized group.

Special instruments already in production by this division include a phase equilibrium computer developed by one of the nation's large oil companies, an automatic titrator for moisture determination developed by a leading pharmaceutical concern, an electrophoresis convection apparatus of interest to the field of biochemistry, Analogue Computer equipment for aircraft companies, and other special-purpose instruments.

Any firm or institution with special instrumentation problems, or with partially developed instrument designs, or with completed prototype instruments already designed and constructed, may secure complete information on this new service by writing Special Products Division, Beckman Instruments, South Pasadena 36, California.

New Labeling Tape

A new smudge-proof pressure-sensitive labeling tape designed for labeling slides, bottles, test tubes, beakers, and other laboratory equipment is announced by the TapeMark Co., 321 Cedar Street, St. Paul, Minn.

The tape's unique writing surface—located beneath a protective covering of transparent acetate film tape—permanently reproduces any message. A stylus, or dry ball-point pen, or pencil can be used to inscribe the message.

Prices and additional details may be had from the TapeMark Co.

Commercial Chemicals

The Matheson Company, Inc. has just announced the publication of a new catalogue, entitled Commercial Chemicals List #1. The list includes 189 of the better known chemicals which are offered in 5 gallon units. This new line bridges the gap between 55 gallon drums available from all manufacturers and Matheson's own line of Fine Organic Chemicals in small put-ups.

Working in close conjunction with the basic manufacturers, Matheson lists the name of the manufacturer along with the manufacturer's own specifications. Prompted by requests of both consumers and manufacturers to supply chemicals in 5 gallon containers from their bulk organic stock, Matheson was already shipping many of these containers prior to publication of this price list. Rather than require consumers to pay the expense of the testing and purification of some of these basic solvents and intermediates Matheson furnishes these industrial materials as received from the suppliers.

Copies of the catalogue may be had by writing The Matheson Company, Inc., East Rutherford, New Jersey.

New Literature

A new bulletin on Eberbach air pressure or vacuum pumps designed for laboratory applications has been published by Eberbach Corporation, Ann Arbor, Michigan. Available upon request, this attractive bulletin gives complete performance and size specifications, suggested uses and illustrated descriptions .

A six-page bulletin describing and illustrating the complete line of Labline sectional laboratory furniture is now available from Labline, Inc., 217 N. Desplaines, Chicago 6, Illinois. The folder explains the Labline system which uses 17 basic standardized metal units to give the service and appearance of custom-built furniture.

Constant Temperature Equipment Catalogue

A new 112-page catalogue on constant temperature equipment has just been published by the Electric Hotpack Co., Inc., 5097 Cottman Avenue, Philadelphia 35, Pennsylvania. The catalogue illustrates and describes the entire line of test and measuring equipment for laboratory and production use. It covers all types of baths for water, oil and Hi-Lo temperature, sterilizers, ovens, both mechanical and gravity, bacteriological incubators, dry and humidified, electric furnaces and portable constant temperature rooms, in addition to temperature and humidity conversion tables. Copies are available on request to the company.

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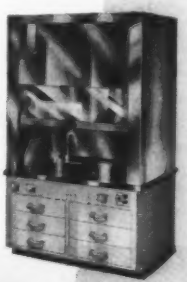

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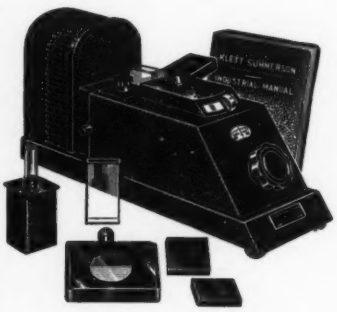

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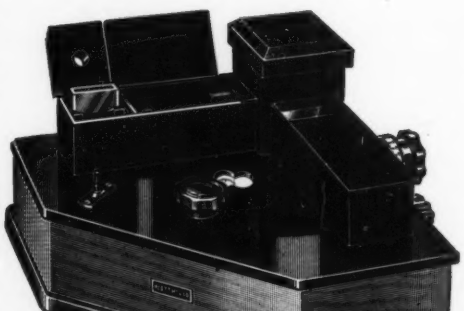

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Incidentally, in Vol. 24, No. 2, of our publication, "Organic Chemical Bulletin," there's an interesting roundup on analytical techniques with the potent complexing reagent Ethylenediaminetetracetic Acid (Eastman P5416). Drop us a line and we'll put you on our mailing list for the Bulletin starting with this issue. No obligation, of course.

For the age of polymers . . .

α,α'-*Azodi-iso-butyronitrile* (Eastman 6400) is a compound perhaps worth considering when you are having difficulty in polymerizing an unsaturated acid. We know little of its potentialities beyond the fact that Roger Adams *et al* found it would polymerize α-acetamidoacrylic acid when ultraviolet light, benzoyl peroxide, etc. had failed (*J.A.C.S.*, 72, 5080, 1950). We also suspect that there is still a certain amount of wealth and fame that can be expected by the chemist who works out a good new polymer from some obscure unsaturated acid—maybe a polymer that makes better golf balls or grommets.

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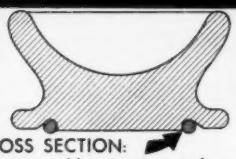

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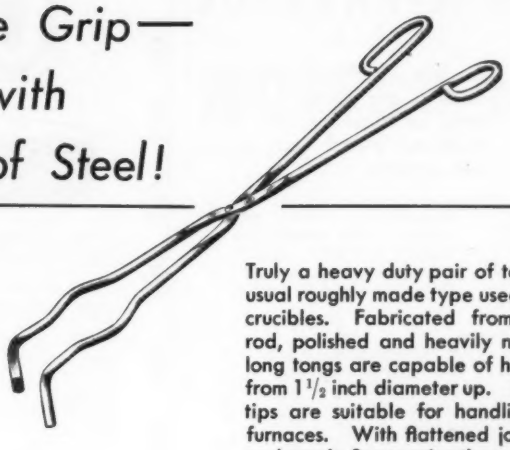

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Amend Drug & Chemical Co., Inc. 30
American Optical Co. 26
Angel & Co., H. Reeve 28
Baker Chemical Co., J. T. ..... IFC
Bersworth Chemical Co. 9
Boekel & Co., Wm. 31
Cambridge Thermionic Corp. 26
Central Scientific Co. 12
Chem Ed Buyers' Guide 34
Coleman & Bell Co., Inc. 32
Coleman Instruments 18
Coors Porcelain Co. 30
Corning Glass Works 14
Delta Chemical Works 33
Eastman Kodak 29
Fisher Scientific ..... BC
Greiner Co., Emil 8
Harshaw Scientific 6
Hevi Duty Electric Co. 19
International Equipment Co. ..... IBC
Interscience Publishers, Inc. 23
Klett Mfg. Co. 28
LaPine & Co., Arthur S. 7
Leeds & Northrup Co. 20
Linde Air Products Co., The (Unit of Union Carbide & Carbon Corp.) 30
McGraw-Hill Book Co., Inc. ..... FC
Macmillan Co., The 10
Mallinckrodt Chemical Works 17
Merck & Co., Inc. 15
Metalab Equipment Corp. 27
New York Laboratory Supply Co., Inc. 24
Palo Laboratory Supplies, Inc. 31
Rupp & Bowman Co. 33
Sargent & Co., E. H. 13
Schleicher & Schuell Co., Carl 24
Standard Scientific Supply Corp. 11
Thomas Co., Arthur H. 3
Union Carbide & Carbon Corp. (The Linde Air Products Co., Unit of) 30
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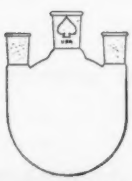

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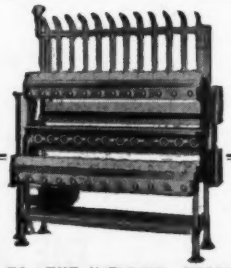

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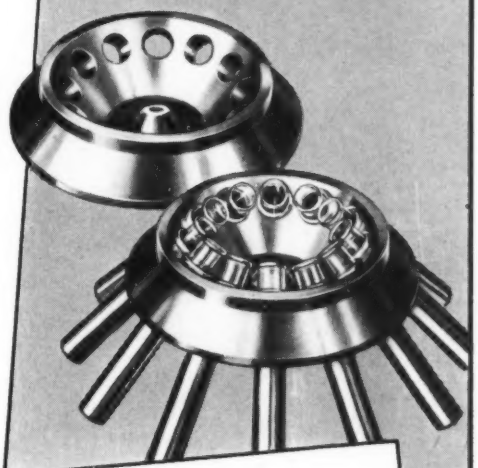

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